CHEM
CHEMISTRY MAIN LIBRARY

BULLETIN

OF THE

CHEMICAL SOCIETY OF JAPAN

OHIO STATE UNIVERSITY
NOV - 1 1960
LIBRARY

Vol. 33 No. 9 September 1960

CONTENTS

General and Physical

M. ICHIKUNI: Sur la Dissolution des Minerais Sulfurés en Divers Milieux. III. Facteurs Intervenant à la Dissolution de la Chalcopyrite 1159

M. ICHIKUNI: Sur la Dissolution des Minerais Sulfurés en Divers Milieux. IV. Dissolution du Minerai Jaune 1162

M. HAYASHI: The Self-diffusion of Surface Active Agents in Nylon 1184

A. AIHARA: Estimation of the Energy of Hydrogen Bonds Formed in Crystals. III. Amides. 1188

Y. IIDA and K. SHIMADA: Effect of Compacting and Grinding on the Hydrogen Reduction of Nickel Oxide 1194

K. FUKUI, H. KATO and T. YONEZAWA: A Molecular Orbital Theory of Saturated Compounds. I. Ionization Potential and Bond Dissociation Energy. 1197

K. FUKUI, H. KATO and T. YONEZAWA: A Molecular Orbital Theory of Saturated Compounds. II. Chemical Reactivity in Nucleophilic Substitution 1201

Y. TANIZAKI and H. ONO: Dichroism of Dyes in the Stretched PVA Sheet. IV. Absorption Spectra of Carbocyanines and Note on the Chemical Structures 1207

M. TAKEDA, K. IIMURA, A. YAMADA and Y. IMAMURA: Studies on the Configurations and Conformation of Polystyrene 1219

Y. KUBOKAWA: Rate of Desorption from a Heterogeneous Surface. 1226

Y. KUBOKAWA: Rate of Adsorption on a Heterogeneous Surface. Determination of the Distribution Function of Sites from Adsorption Rate Curves. 1227

J. HIGUCHI, S. SHIDA, R. KUSAKA and T. MIYAMAE: Decay of Free Radicals in γ-Irradiated Polymethyl Methacrylate 1232

K. HIROMI: Theory of the Influence of the Dielectric Constant on the Rate of Reaction in Solution with Application to Enzyme Reactions. I. Development of the Theory and its Application to Some Simple Systems 1251

K. HIROMI: Theory of the Influence of the Dielectric Constant on the Rate of Reaction in Solution with Application to Enzyme Reactions. II. Application of the Theory to Enzyme Reactions 1264

T. KUGE and S. ONO: Amylose-Iodine Complex. II. Effects of Inorganic Salts. 1269

T. KUGE and S. ONO: Amylose-Iodine Complex. III. Potentiometric and Spectrophotometric Studies 1273

Analytical and Inorganic

S. KIDA: Investigation on Mixed Complexes. IV. The Formation of a Mixed Complex by Reaction of Tetracyanoniccolate(II) with Bis-(dithiooxalato)-niccolate(II) 1204

N. TANAKA and K. KATO: The Kinetics of Substitution Reactions Involving Metal Complexes. IV. Reaction between Lead and Zinc-Ethylenediaminetetraacetate Complexes in Aqueous Solutions 1236

(Continued on inside cover)

Published by the Chemical Society of Japan

THE CHEMICAL SOCIETY OF JAPAN
(NIPPON KAGAKUKAI)

Founded in 1948

Successor to the former Chemical Society of Japan founded in 1878 and to the Society of Chemical Industry, Japan founded in 1898.

(*Continued from outside cover*)

Y. MUTO: Metal Complexes with Terdentate Ligands. II. Synthesis and Properties of Tri-coordinated Copper(II) Complexes 1242

S. YAMADA, H. NISHIKAWA and R. TSUCHIDA: Spectrochemical Study of Microscopic Crystals. XXII. The Structure of Cupric Mono-, Di- and Trichloroacetate 1278

Organic and Biological

T. MATSUURA, K. SAITO and Y. SHIMAKAWA: Oxidation of Terpene Compounds with *tert*-Butyl Chromate. VI. Oxidation of Terpinolene 1151

S. S. MUHAMMAD: Thermal Decarboxylation of *p*-Chlorobenzoic Acid. 1154

T. MUROI: Synthesis of 2-Phenyltropone and its Bromo-derivatives from 2-Phenylsuberone. . .1166

S. MURAHASHI, S. NOZAKURA and M. SUMI: Stereoregular Polymers. VI. Copolymerization with the Ziegler Catalysts 1170

T. AMIYA: The Aconite Alkaloids. XXXIV. On Lucidusculine. Part IV. 1175

N. TOKURA, R. TADA and R. IGARASHI: Halogen Substitution in Liquid Sulfur Dioxide. I. Formation of Alkyl Halides from Alcohols and Halogens 1176

T. TACHIKAWA: On the Oxidation of β-Aminopropionitrile. I 1181

K. FUKUI, K. MOROKUMA and C. NAGATA: A Molecular Orbital Treatment of Phosphate Bonds of Biochemical Interest. I. Simple LCAO MO Treatment 1214

I. SEKIKAWA: Oxidation of 5-Methyl-2,1,3-benzothiadiazole with Potassium Permanganate. . .1229

T. MUKAI: Structure of Namakochrome. II 1234

Y. KAWASE, K. OGAWA, S. MIYOSHI and K. FUKUI: Synthetic Studies on the Benzofuran Derivatives. VI. Synthesis of Furano(2'',3'':7,8)-2'-methoxy-isoflavone Derivatives. 1240

T. NOZOE, T. MUKAI, T. NAGASE and Y. TOYOOKA: Tropones. V. Addition Reaction of Tropone and Maleic Anhydride 1247

T. NOZOE, K. DOI and S. ENDO: The Reaction Products of 2,4,7-Tribromotropone and Anilines. .1285

F. TODA and M. NAKAGAWA: Further Studies on the Intramolecular Cyclization of *o,o'*-Dihydroxydiphenyldiacetylene 1287

S. AKIYAMA and M. NAKAGAWA: Cyclic Acetylenes. III. Synthesis of 1,1'-Dianthryldiacetylene. A Model Substance of a Cyclic Tetraacetylene Containing Anthracene Nuclei 1291

S. AKIYAMA, S. MISUMI and M. NAKAGAWA: Cyclic Acetylenes. IV. Synthesis of a Cyclic Tetraacetylene Containing Anthracene Nuclei 1293

K. YAMADA, H. NAKATA and Y. HIRATA: The Structure of Aureothin. X. Aza Derivatives of Aureothin, Isoaureothin and Desmethylisoaureothin 1298

K. YAMADA, T. NAITO, K. OKUHARA, H. NAKATA and Y. HIRATA: The Structure of Aureothin. XI. On the Structure of Some of the Isoaureothin Derivatives 1303

Short Communications

A. TERAHARA, M. OHASHI, K. NAKANISHI, I. YAMAGUCHI and N. HAYAKAWA: NMR Spectra of the Metabolites of Monascus. I. NMR Spectra of Several Authentic Pyrones. 1310

M. OHASHI, A. TERAHARA and K. NAKANISHI: NMR Spectra of the Metabolites of Monascus. II. Monascorubrin 1311

M. OHASHI, A. TERAHARA, K. NAKANISHI, I. YAMAGUCHI and N. HAYAKAWA: NMR Spectra of the Metabolites of Monascus. III. Monascofalvin and Reduction Products 1312

T. SHINODA, H. SUGA and S. SEKI: On the Unusual Pre-melting Phenomenon of KSCN Crystal. .1314

K. SAITO and M. TSUCHIMOTO (née TAMURA): Isotopic Exchange of Gallium between the Gallium(III) Ion and its Ethylenediamine'*N*, *N*, *N'*, *N'*-tetraacetate and *N*-2-Hydroxyethyl-ethylenediamine-*N*, *N'*, *N'*-triacetate in an Aqueous Solution 1315

K. HIROTA, G. MESHITSUKA, F. TAKEMURA and T. TANAKA: The Polymerization of α-Methylstyrene by Iodine Catalyst 1316

T. WADA, E. KISHIDA, Y. TOMIIE, H. SUGA, S. SEKI and I. NITTA: Crystal Structure and Thermodynamical Investigations of Triethylenediamine, $C_6H_{12}N_2$. 1317

Oxidation of Terpene Compounds with tert-Butyl Chromate. VI. Oxidation of Terpinolene

By Tamon MATSUURA, Kozo SAITO and Yasuo SHIMAKAWA

(Received February 4, 1960)

On oxidation of terpene hydrocarbons with *tert*-butyl chromate[1-5] it was recognized that the active methylene group adjacent to a double bond in a molecule was mainly oxidized to α, β-unsaturated carbonyl. Among diene terpene hydrocarbons, only D-limonene has been researched and the oxidation products were rather more complicated than in the cases of α-pinene[1,5], D-carvomenthene[2,5] and D, L-α-terpinyl acetate[4]; D-limonene was oxidized to D, L-carvone, D-isopiperitenone, piperitenone and carvacrol. This report is concerned with the oxidation of terpinolene (I) which has an active methylene group at $C_{(3)}$ position situated between two double bonds.

Results and Discussion

The oxidizing agent was prepared in a manner similar to that in the previous reports[3,4] from chromic anhydride and *tert*-butyl alcohol, glacial acetic acid and acetic anhydride being added.

Experiment 1.—The mole ratio of sample to chromic anhydride was 1 to 2. After the reaction was carried out at 10°C during 15 hr., the reaction mixture which resulted was treated with water and oxalic acid under ice-cooling to decompose the excess of the oxidant. Then the mixture was separated into a benzene and an aqueous layer. From the benzene layer was obtained a neutral brown oil which was fractionated and the constituents of which were identified.

Experiment 2.—The mole ratio of sample to chromic anhydride was 1 to 2, and the reaction was carried out at 10°C during 45 hr. The reaction mixture was treated in the same manner as in Experiment 1.

Experiment 3.—In this experiment glacial acetic acid and acetic anhydride were excluded from the oxidant. This is not the case in Experiments 1 and 2. When the oxidant was added to the sample solution at 0°C during 2 hr., the reaction was stopped and the reaction products were examined.

In the case of each of the above three experiments, piperitenone (II), 8-hydroxy-*p*-cymene (III), *p*-cymene (IV) and α, p-dimethylstyrene (V) were identified as constituents of the reaction products, and an unidentified carbonyl compound was also recognized.

Experiment 4.—*p*-Cymene was treated with *tert*-butyl chromate under the same condition as in Experiment 1, and small amounts of 8-hydroxy-*p*-cymene and *p*-methylacetophenone were identified as reaction products.

The above results are summarized in Table I.

In the oxidation of terpinolene with *tert*-butyl chromate, the point different from the oxidation of the other cyclic terpene compounds already reported is the aromatization of the hydroaromatic ring by dehydrogenation in addition to the oxidation of the active methylene group adjacent to a double bond. On oxidation of D-limonene and D, L-α-terpineol it has been clarified that among two active methylene

TABLE I

No. of Experiment	Sample	Reaction condition: Sample : CrO_3 mole ratio	CH_3COOH wt. % to sample	$(CH_3CO)_2O$ wt. % to sample	Temp. °C	Time hr.	Reaction product wt. % to sample:				
1	Terpinolene	1 : 2	17	17	10	15	6	8	1	2	0
2	Terpinolene	1 : 2	50	50	10	45	11	6	2	3	0
3	Terpinolene	1 : 2	0	0	0	2	4	4	2	1	0
4	*p*-Cymene	1 : 2	17	17	25	15	0	0.8	—	0	0.2

1) T. Matsuura and K. Fujita, *J. Sci. Hiroshima Univ.*, **A16**, 173 (1952).
2) K. Fujita and T. Matsuura, ibid., **18**, 455 (1955).
3) K. Fujita, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **78**, 1112 (1957).
4) T. Matsuura, T. Suga and K. Suga, ibid., **78**, 1122 (1957).
5) G. Dupont, R. Dulou and O. Mondou, *Bull. soc. chim. France*, **1953**, 60.

groups adjacent to a double bond in the carbon ring the $C_{(6)}$ active methylene is predominantly oxidized in comparison with the $C_{(3)}$ active methylene. In the case of terpinolene, however, it was found that the oxidation of the $C_{(3)}$ active methylene has proceeded in exclusion of the oxidation of the $C_{(6)}$ active methylene, owing to the existence of a double bond between $C_{(4)}$ and $C_{(8)}$ atoms.

The formation of piperitenone II from terpinolene I by the oxidation of the $C_{(3)}$ active methylene is an analogous reaction to the formation of α,β-unsaturated ketone from cyclic terpene compounds such as α-pinene, limonene, carvomenthene and α-terpinyl acetate reported already. However the formation of 8-hydroxy-*p*-cymene III and *p*-cymene IV from terpinolene is due to oxidation of the $C_{(8)}$ position and dehydrogenation, which was found for the first time in the oxidation of terpinolene. An exocyclic double bond between $C_{(4)}$ and $C_{(8)}$ carbons in terpinolene seems to migrate into the carbon ring in the presence of acid to form α-terpinene[6]. It has been reported that menthene-4(8) is rearranged into menthene-3 by the action of dilute sulfuric acid solution[7]. Thus it may be considered that an exocyclic double bond of terpinolene molecule migrates into the ring to form a conjugated diene (α-terpinene) by acid reagent. Although α-terpinene was not found in the reaction products, it has reacted easily with *tert*-butyl chromate to produce a large amount of *p*-cymene[*1]. Therefore *p*-cymene in the reaction products is considered to be produced by dehydrogenation of α-terpinene which results from isomerization of terpinolene.

It would seem that 8-hydroxy-*p*-cymene has been formed from *p*-cymene by the oxidation of an active methine group at $C_{(8)}$. But this is not the case, because the result of the oxidation of *p*-cymene with *tert*-butyl chromate in the same conditions as in Experiment 1 (sample, terpinolene) has shown that *p*-cymene is virtually stable against *tert*-butyl chromate, giving only a small amount (1% to the sample) of 8-hydroxy-*p*-cymene and of *p*-methylacetophenone.

Accordingly it can not be considered that 8-hydroxy-*p*-cymene is formed via *p*-cymene.

It has been recognized that on distillation of 8-hydroxy-*p*-cymene in presence of acid (a small amount) α, *p*-dimethylstyrene is produced by dehydration[8], so α, *p*-dimethylstyrene is considered to be a secondary product formed by dehydration during the oxidation reaction or distillation of the reaction products.

On the basis of the above results and discussion the reaction course is summarized as shown in Fig. 1.

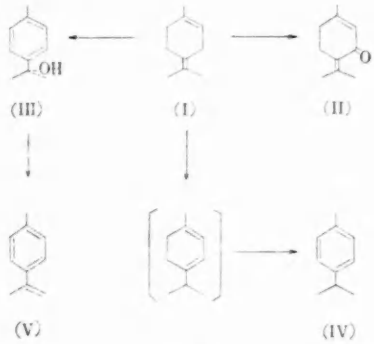


Fig. 1

Experimental

1. Oxidation of Terpinolene

1.1. Preparation of Terpinolene

Terpinolene (b. p.$_{21}$ 78.0°C, n_D^{25} 1.4859, d_4^{25} 0.8577, tetrabromide m. p. 116°C) was prepared, according to Charalton and Day's method[9] by treating α-pinene (b. p.$_{28}$ 58~59°C, n_D^{25} 1.4628, d_4^{25} 0.8548, $[\alpha]_D$ +40.62°) with 50% alcoholic solution of sulfuric acid, and by rectifying the crude product.

1.2. Oxidation and Identification of Reaction Products

1.2.1. Experiment 1

tert-Butyl chromate was prepared by adding 88 g. of chromic anhydride in small portions to 163 g. of *tert*-butyl alcohol under ice-cooling, diluting the resultant with 650 g. of benzene, and drying the resulting benzene solution on anhydrous sodium sulfate. To the solution were added 10 g. of glacial acetic acid and 10 g. of acetic anhydride just prior to the oxidation. This oxidant was added dropwise with stirring at 10°C to 60 g. of terpinolene in 60 g. of dry benzene during 1 hr., and the mixture was left to stand at 10°C for 15 hr. Then 700 cc. of water and 160 g. of oxalic acid were added under ice-cooling to the viscous reaction mixture and shaken to decompose the excess of the oxidant. The mixture was separated into a benzene layer and an aqueous one, and the latter was extracted with benzene. The combined benzene solution was washed with 10% aqueous solution of sodium carbonate and with water successively, and dried (anhydrous sodium sulfate). After removal of the solvent, 28.1 g. of a neutral brown oil was obtained.

The aqueous solution was neutralized with a sodium carbonate solution and extracted with chloroform. From the extract 7.9 g. of a dark-brown viscous oil was obtained. The sodium carbonate solution was acidified with 10% sulfuric acid solution and

6) O. Wallach, *Ann.*, **275**, 103 (1893).
7) O. Wallach, *Ber.*, **39**, 2504 (1906); *Ann.*, **360**, 70 (1908).
*1) T. Matsuura, unpublished.
8) J. N. Borglin, D. A. Lister, E. J. Lorand and J. E. Reese, *J. Am. Chem. Soc.*, **72**, 4591 (1950).
9) R. W. Charalton and A. R. Day, *Ind. Eng. Chem.*, **29**, 92 (1937).

extracted with chloroform, giving 18.9 g. of a dark-brown viscous oil. As neither of the above dark-brown viscous oils could be distilled under reduced pressure and also could not be crystallized, they seemed to be polymer, the further test of which was not undertaken.

The neutral oil was fractionated, giving the physical properties of Fig. 2, and then the constituents of fractions were identified.

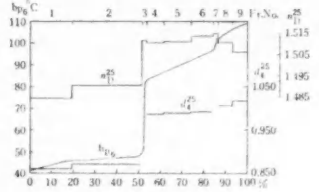


Fig. 2

(i) *p*-Cymene and α,p-Dimethylstyrene. Fractions 1 and 2 gave a tetrabromide of m. p. 116°C which was identified as terpinolene tetrabromide by admixture with an authentic sample. On the other hand, the ultraviolet absorption spectra of these two fractions showed the absorption maxima at 247 and 273 mμ which coincide with those of α,p-dimethylstyrene (247 mμ)[10] and *p*-cymene (273 mμ)[11].

Ozone was passed in a solution of 5 g. of the combined fractions 1 and 2 dissolved in 20 cc. of ethyl acetate-glacial acetic acid (6 : 2) under ice-cooling. The ozonized solution, after the solvent having been driven off under reduced pressure and 50 cc. of water being added, was steam-distilled to decompose the ozonide. The distillate was extracted with ether, and the ether extract was washed with 10% aqueous sodium carbonate solution and dried. From the ether extract thus treated, 2.1 g. of an oily product was obtained, after the solvent was driven off. The oily product was eluted through a silica gel column with petroleum benzine and ether successively, and the following two fractions, A and B, were obtained.

Fraction	Yield (g.)	λ_{max} (mμ) (absorptivity a*2)
A	1.20	254(1.42), 260(2.35), 265(2.86), 267(2.95), 273(3.34)
B	0.67	252(68.90)

Fraction A (b. p.$_{15}$ 62~64°C, n_D^{25} 1.4795, d_4^{25} 0.8552) was confirmed to be *p*-cymene from its ultraviolet spectrum. Ultraviolet absorption spectrum of fraction B coincided with that of *p*-methylacetophenone (λ_{max}; 252 mμ)[12]. 2,4-Dinitrophenylhydrazone m. p. 256.5~257.0°C and semicarbazone m. p. 205~206°C of fraction B were identified to be those of *p*-methylacetophenone by mixing with authentic samples. The formation of *p*-methylacetophenone by ozonolysis of fractions 1 and 2 shows the existence of α,p-dimethylstyrene. From the result of ozonolysis the amounts of *p*-cymene and α,p-dimethylstyrene in fractions 1 and 2 (unified) were calculated to be 24% and 13% respectively.

(ii) 8-Hydroxy-*p*-cymene. Infrared absorption spectra of fractions 4 and 5 showed a hydroxy band in a region of 3400 cm^{-1}, and from these fractions a phenylurethane of m. p. 118~119°C was obtained, which was identified as that of 8-hydroxy-*p*-cymene*3 by the mixed melting point determination.

Three grams of the combined fractions 4 and 5 was heated with 3 g. of powdered potassium hydrogen sulfate at 190~200°C and distilled. The distillate was extracted with petroleum benzine, and the extract was washed with 10% aqueous sodium carbonate solution, dried and eluted through a silica gel column. From the extract thus treated 0.72 g. of an oily product was obtained after the solvent was driven off. Its infrared spectrum showed absorption bands at 1630 cm^{-1} (conjugated double bond), 883 cm^{-1} (end methylene) and 816 cm^{-1} (aromatic *para* substitution), and its ultraviolet spectrum had an absorption maximum at 246 mμ. Thus the dehydration product was confirmed to be α,p-dimethylstyrene, and therefore the original alcohol to be 8-hydroxy-*p*-cymene.

(iii) Piperitenone. Ultraviolet absorption spectra of fractions 6 and 7 showed λ_{max} 242 mμ (ε : 10140) and 276 mμ (ε : 7640) which coincide with absorption maxima of piperitenone, 242 mμ (ε : 12500) and 278 mμ (ε : 7920)[13]. 2,4-Dinitrophenylhydrazone of the above fractions melted at 183.5~184°C. This was identified as that of piperitenone by mixing with a known sample.

Fractions 8 and 9 were unified and eluted through a silica gel—plaster of paris (4 : 1) column with ether-petroleum benzine (3 : 7), and it was verified that 37% of the above unified two fractions was piperitenone.

(iv) Unidentified Carbonyl Compound. By column chromatography of fractions 8 and 9 mentioned above, a carbonyl compound was separated, which gave a 2,4-dinitrophenylhydrazone of m. p. 216.5~217°C and ultraviolet absorption spectrum which had λ_{max} 244 mμ and 257 mμ, but it was not identified.

1.2.2. Experiment 2

tert-Butyl chromate solution was prepared, in the same way as in Experiment 1, from chromic anhydride (147 g.), *tert*-butyl alcohol (274 g.), benzene (1000 g.), glacial acetic acid (50 g.), and acetic anhydride (50 g.). This oxidant was added dropwise to 100 g. of terpinolene in 100 g. of dry benzene, and the reaction was carried out for 45 hr. at 10°C. Then the reaction mixture was treated as in Experiment 1, and 56.1 g. of a neutral oil was obtained. Piperitenone, 8-hydroxy-*p*-cymene, α,p-dimethylstyrene and *p*-cymene were found in the neutral oil*4 (cf. Table I).

10) M. J. Murray and W. S. Gallaway, *J. Am. Chem. Soc.*, **70**, 3867 (1948).
11) R. T. O'Connor and L. A. Goldblatt, *Anal. Chem.*, **26**, 1726 (1954).
12) J. Tanaka, S. Nagakura and M. Kobayashi, *J. Chem. Phys.*, **24**, 311 (1956).
*2 The concentration is in g./l. and the cell length in cm.
*3 8-Hydroxy-*p*-cymene was prepared from *p*-bromotoluene and acetone by Grignard reaction (V. N. Ipatieff, H. Pines and R. C. Obberg, *J. Am. Chem. Soc.*, **70**, 2123 (1948)).
13) R. Y. Naves, *Bull. soc. chim. France*, **1951**, 370.
*4 Identification of the constituents in the neutral oils in 1. 2. 2 and 1. 2. 3 was carried out by the similar methods as in 1. 2. 1.

1.2.3. Experiment 3

tert-Butyl chromate solution was prepared from the same amounts of chromic anhydride, *tert*-butyl alcohol and benzene as in Experiment 1, but glacial acetic acid and acetic anhydride were excluded. Sixty grams of terpinolene in 60 g. of dry benzene was stirred with the above oxidant for 2 hr. at 0°C. Then the reaction mixture was treated as in Experiment 1, and 28.5 g. of a neutral oil was obtained. The same compounds as in Experiment 2 were confirmed to exist in the neutral oil (cf. Table I).

2. Oxidation of *p*-Cymene (Experiment 4)

To 60 g. of *p*-cymene (b.p.$_{22}$ 70.5°C, n_D^{25} 1.4860, d_4^{25} 0.8519) dissolved in benzene (60 g.) was added dropwise an oxidant solution consisting of chromic anhydride (89 g.), *tert*-butyl alcohol (166 g.), benzene (650 g.), glacial acetic acid (10 g.) and acetic anhydride (10 g.), and the reaction was carried out for 15 hr. at 25°C. Then the reaction mixture, after being treated as in the preceding experiments, gave 56.4 g. of a neutral oil which was fractionated as follows: fraction 1 (53 g., b.p.$_{25}$ 73~74°C, n_D^{25} 1.4876, d_4^{25} 0.8645), fraction 2 (0.6 g., b.p.$_4$ 46~70°C, n_D^{25} 1.5017) and residue (0.4 g.).

It was confirmed that fraction 1 was unchanged *p*-cymene by its ultraviolet spectrum measurement. Fraction 2 gave 2,4-dinitrophenylhydrazone m.p. 256°C of *p*-methylacetophenone and phenylurethane m. p. 120°C of 8-hydroxy-*p*-cymene, both of which were identified by mixing with authentic samples. An ultraviolet absorption spectrum of fraction 2 have λ_{max} 250 mμ (a: 20) which agrees virtually with λ_{max} 252 mμ (a: 117) of *p*-methylacetophenone.

Summary

Oxidation of terpinolene with *tert*-butyl chromate was carried out, and piperitenone, 8-hydroxy-*p*-cymene, *p*-cymene and α,*p*-dimethylstyrene were obtained as reaction products. In this oxidation it was found that the active methylene group at $C_{(3)}$ position in terpinolene was oxidized to a carbonyl group, and that the aromatization of carbon ring in terpinolene molecule took place. This could not be recognized in the oxidation of terpene compounds already reported by this oxidant.

This research was aided by a grant from the Ministry of Education.

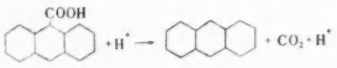
Department of Chemistry
Faculty of Science
Hiroshima University
Hiroshima

Thermal Decarboxylation of p-Chlorobenzoic Acid

By S. S. MUHAMMAD

(Received February 13, 1960)

Though Fieser and Fieser[1a] had suggested in 1944 the probability of an electrophilic bimolecular (S_E2) mechanism for the decarboxylation of *o*- and *p*-hydroxybenzoic acids, no experimental evidence was available for this till 1948, when Schenkel and Schenkel-Rudin[2] recognized this as likely for the decomposition of anthracene-9-carboxylic acid:



since Schenkel[3] had established the fact that this acid underwent decarboxylation more readily in acidic solvents such as sulfuric or dichloroacetic acid than in basic solvents such as 7:8 benzoquinoline. Schubert[4] tried to establish a relation between the rate of decarboxylation and hydrogen ion concentration of the medium, by studying the decomposition of mesitoic acid in aqueous sulfuric acid of 80 to 100%. Johnson and Heinz[5] obtained evidence for the S_E2 mechanism by studying the rates of decarboxylation, under identical conditions, of different substituted cinnamic acids, $PhCR'=CR''COOH$. The order of rates found by them agreed with the electronic effects of the substituents R′ and R″ on the basis of the following bimolecular mechanism:

$$PhCR'=CR''COOH + HA \rightleftharpoons \overset{+}{P}hCR'CHR''COOH + A'$$

followed by a comparatively fast step:

$$\overset{+}{P}hCR'CHR''COOH \rightarrow PhCR'=CHR''CO_2 + H^+$$

1) L. Fieser and M. Fieser, "Organic Chemistry", Heath & Co., Boston 1944; a p. 674, b 568.
2) H. Schenkel and M. Schenkel-Rudin, *Helv. Chim. Acta*, **31**, 514 (1948).
3) H. Schenkel, ibid., **29**, 436 (1946).
4) W. M. Schubert, *J. Am. Chem. Soc.*, **71**, 2639 (1949).
5) W. S. Johnson and W. E. Heinz, ibid., **71**, 2913 (1949).

Then Brown, Hammick and Scholefield[6] showed that the decarboxylation of salicylic, 2, 4-dihydroxy- and 2, 4, 6-trihydroxybenzoic acids occurs by an S_E2 mechanism. These studies were extended by S. S. Muhammad and Mrs. Aziz Mirza on benzoic acid[8], *o*-chloro- and 2, 4-dichlorobenzoic acids[9,10]. The technique employed was primarily to collect the carbon dioxide evolved in a gas burette. But in the present study a slightly different technique has been adopted. The results reported pertain to *p*-chlorobenzoic acid and these will be shortly extended to other acids.

Experimental and Results

Materials. — All the chemicals were of analar grade, and were used without further purification. The solvents, resorcinol, glycerol and aniline were of extra pure quality and when heated alone did not emit carbon dioxide.

Apparatus.—The heating device consisted of an oil bath heated electrically, the temperature control being achieved by means of a Sunvic energy regulator and a Sunvic relay. Temperatures were recorded with a 300°C thermometer having an accuracy of ±0.5°C. The bath liquid was a local sweet oil which, on heating for 24 hr. at 300°C, emitted volatile vapours and then behaved as a nonvolatile liquid below this temperature.

The reaction vessel was a modified Walton's[11] apparatus, the mouth of which was made of ground glass joint with stopper to which a glass tube was attached and this served as delivery tube for carbon dioxide liberated. The side tube had also a neck made of glass joint. The stopper for this had a spoon device where a weighed amount of the acid could be kept in a small test tube, which could be dropped by a slight turn of the stopper. Both the stoppers were lubricated with the least quantity of Silicone grease.

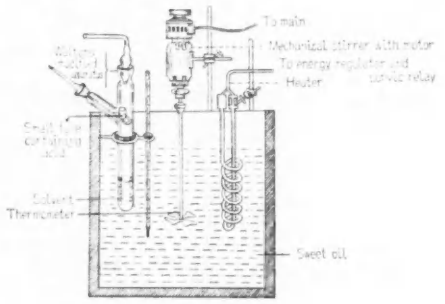


Fig. 1.

The carbon dioxide was absorbed in baryta solution kept in an Erlenmeyer flask.

The oil thermostat with the modified Walton apparatus is shown in Fig. 1.

Procedure.—In order to find the order of reaction, different amounts of *p*-chlorobenzoic acid were taken with 20 g. of resorcinol in the Walton tube and heated for a definite interval. The liberated carbon dioxide was absorbed in baryta solution of approximately N/10 strength. Its exact titre value was determined by standardized N/10 oxalic acid before and after the absorption of carbon dioxide. The amount of acid decomposed, x, was calculated by the relation: $\Delta v \times 10^{-4}$ g. mol., where Δv is the difference in the titre values 'v_0' initially and 'v' after absorption. Since the amount of acid initially taken, 'a', is known, the amount of acid remaining at any interval, '$a-x$', can be found.

In the first series of experiments, the initial amounts 'a' were converted into relative concentrations, 'c', and the corresponding amounts 'x' decomposed in the same interval were converted into relative rates 'r' (Table I).

The plot of log c against log r yields a slope of 1, indicating the order to be one.

In the second series of experiments, the same initial amount of the acid 'a', was heated. The carbon dioxide evolved was absorbed in the same volume of the baryta solution of the same concentration. The latter solution was titrated by N/10 oxalic acid after a desired interval. The same operation being repeated for different desired intervals. Table II contains these data.

The plot of log $(a-x)$ against time t gives straight line proving the first order character of the reaction.

The rate constants can be obtained by the equation $k=2.303/t \cdot \log a/a-x$. Thus using the data of Expt. 4, we get $k=2.303/12 \log 29.4/19.6 = 3.378\times10^{-2}\,\mathrm{min}^{-1} = 5.63\times10^{-4}\,\mathrm{sec}^{-1}$. Alternatively the rate constant can be derived by the relation

$$k' = \frac{x}{t} \times \frac{1}{\bar{a}}$$

where x is the amount decomposed in the interval t and $\bar{a}$ is the average concentration during this interval i. e. $\{a+(a-x)\}/2$. Thus the data of Expt. 4, give

$$k' = \frac{9.8\times10^{-4}}{12} \times \frac{1}{(29.4+19.6)\times10^{-4}/2}$$

$$= \frac{9.8\times10^{-4}}{12\times24.5\times10^{-4}} = 3.33\times10^{-2}\,\mathrm{min}^{-1}$$

$$= 5.56\times10^{-4}\,\mathrm{sec}^{-1}$$

In subsequent experiments, the latter method of deriving the rate constant has been used.

In the third series of experiments, different initial amounts of the acid, 'a', were heated at constant temperature, T, and the amounts 'x', decomposed in 12 minutes were determined in each case. The rate constants were calculated as above. At each temperature, 5 observations were made and the mean of these was taken as the rate constant at that temperature. These results are presented in Table III.

6) B. R. Brown, D. Ll. Hammick and A. J. B. Scholefield, *J. Chem. Soc.*, **1950**, 778.
7) B. R. Brown, W. W. Elliott and D. Ll. Hammick, ibid., **1951**, 1384.
8) S. S. Muhammad and Aziz Siddiqui, *J. Indian Chem. Soc.*, **31**, 726 (1954).
9) S. S. Muhammad and Aziz Siddiqui, ibid., **33**, 570 (1956).
10) Aziz Mirza, Ph. D. thesis approved by the Osmania University, 1957 pp. 100—106.
11) J. H. Walton, *Physik. Chem.*, **47**, 185 (1904).

TABLE I. SOLVENT RESORCINOL, $T=235°C$

Weight of acid, g.	Weight, g., decomposed per sec.	Relative concentration c	Relative rate r	$\log c$	$\log r$
0.12	5.83×10^{-5}	1	1	0	0
0.16	8.00	1.33	1.37	0.1239	0.1367
0.20	10.00	1.67	1.72	0.2227	0.2355
0.25	12.17	2.08	2.09	0.3181	0.3201
0.35	17.5	2.92	3.00	0.4654	0.4771
0.40	20.33	3.33	3.49	0.5224	0.5428
0.50	25.67	4.17	4.40	0.6201	0.6435
0.60	29.67	5.00	5.09	0.6990	0.7067

TABLE II.

A 0.460 g. portion of acid was taken in each experiment (i.e. $a=29.4\times10^{-4}$ g. mol.). The carbon dioxide formed was absorbed in 40 cc. of baryta solution which was titrated with N/10 oxalic acid.

Expt. No.	Time, t min.	Titre value of N/10 Oxalic acid, T	Amount of acid decomposed, x ($x=T\times10^4$ g. mol.)	$(a-x)\times10^4$	$\log(a-x)$
1	0	33.4	—	29.4	1.4683
2	4	29.8	3.6	25.8	1.4116
3	8	26.4	7.0	22.4	1.3502
4	12	23.6	9.8	19.6	1.2923
5	15	21.4	12.0	17.4	1.2405
6	20	19.0	14.4	15.0	1.1761
7	26	16.5	16.9	12.5	1.0969
8	32	15.0	18.4	11.0	1.0414

In order to examine the above results, it was desirable to derive the activation energy E, frequency factor A, free energy of activation ΔF° and entropy of activation ΔS°.

The plot of $\log k$ vs. $1/T$ (Fig. 2) gave the activation energy E.

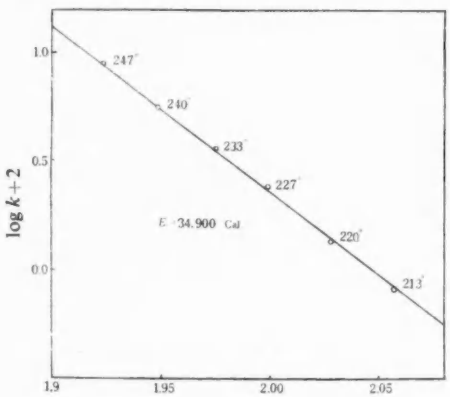


Fig. 2.

The Arrhenius equation $k=Ae^{-E/RT}$ was used to calculate the frequency factor A and the Eyring equations[12]

$$\Delta F^\circ=2.303\ RT\ \log\ (RT/Nhk)$$

$$\Delta S^\circ=(\Delta H^\circ-\Delta F^\circ)/T$$

yielded the free energy and entropy of activation. The term ΔH° is the enthalpy of activation, taken equal to $E-RT$, where E is the observed activation energy and T the temperature at which entropy and free energy are being calculated. The values of the constants used are[13]: $R=1.987$ cal. and 8.31439×10^7 erg, $N=6.02308\times10^{23}$, $h=6.6252\times10^{-27}$ erg, $T^0=$ °C+273.15. The results of these calculations are shown in Table IV.

Effect of Other Solvents.—The solvent glycerol is also effective in bringing about decarboxylation of *p*-chlorobenzoic acid, but the rate of decarboxylation is less in glycerol than in resorcinol. Thus, while at 233°C in resorcinol 2.5% of the acid decomposed per minute, in glycerol under identical conditions the rate of decarboxylation is 1.9% per minute. Besides, the rate constants for decarboxylation in glycerol calculated from the first and the second order rate equations show large trends.

Addition of a small quantity of aniline to glycerol solution retards the rate of decarboxylation and when enough of aniline is added, the reaction completely stops. When the acid was dissolved in pure aniline, it did not decompose at 247°C, which was the highest temperature employed in the present investigation.

12) S. Glasstone, K. J. Laidler and H. Eyring, "The Theory of Rate Processes", McGraw Hill Book Co., (1941).

13) G. W. C. Kaye and T. H. Laby, "Tables of Physical and Chemical Constants", Longmans (1956).

TABLE III.

Expt. No.	$a\times10^4$ g. mol.	$x\times10^4$ g. mol.	$a\times10^4$ g. mol.	$k=\left(\frac{x}{a}\times\frac{1}{12}\right)\text{min}^{-1}$
		(a) $T=213^\circ$C		
1	58.8	5.6	56.0	0.83×10^{-2}
2	44.1	4.3	41.95	0.85
3	39.2	3.6	37.4	0.80
4	35.3	3.4	33.6	0.84
5	29.4	2.9	27.95	0.86
			Mean	0.84
		(b) $T=220^\circ$C		
1	58.8	8.6	54.5	1.32
2	44.1	6.6	40.8	1.35
3	39.2	6.0	36.2	1.38
4	35.3	5.2	32.7	1.33
5	29.4	4.6	27.1	1.42
			Mean	1.36
		(c) $T=227^\circ$C		
1	73.5	18.2	64.4	2.36
2	58.8	14.3	51.4	2.40
3	44.1	11.2	38.5	2.42
4	35.3	8.6	31.0	2.31
5	23.5	6.0	20.5	2.44
			Mean	2.39
		(d) $T=233^\circ$C		
1	73.5	25.6	60.7	3.51
2	58.8	20.8	48.4	3.58
3	44.1	16.0	36.1	3.69
4	35.3	12.2	29.2	3.48
5	23.5	8.4	19.3	3.63
			Mean	3.58
		(e) $T=240^\circ$C		
1	88.2	43.8	66.3	5.51
2	70.6	35.0	53.1	5.49
3	58.8	29.4	44.1	5.55
4	44.1	22.1	33.05	5.57
5	29.4	14.8	22.0	5.61
			Mean	5.55
		(f) $T=247^\circ$C		
1	88.2	61.6	57.4	8.94
2	70.6	49.4	45.9	8.97
3	58.8	42.0	37.8	9.26
4	44.1	30.8	29.0	8.68
5	29.4	20.4	19.2	8.86
			Mean	8.94

TABLE IV.

Property	k, 220°C	E	$\log A$	ΔH 220°C	ΔF°	ΔS° 220°C
Value	2.25×10^{-4} sec^{-1}	34900 cal.	11.81	33900 cal.	37600 cal.	-7.5 cal. deg^{-1}

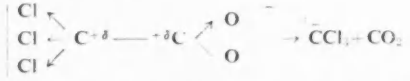
Discussion

The data presented above show that *p*-chlorobenzoic acid is stable in basic solvents and is unstable in acidic solvents. In this respect, it is analogous with salicylic acid[6] and orthochlorobenzoic acid[9] which undergo decomposition through an electrophilic bimolecular (S_E2) mechanism, involving the replacement of the carboxyl:

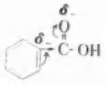
$$H^+ + RC(=O)-O-H \rightarrow RCH + C(=O)-\overset{\oplus}{O}-H$$

or $RCH + CO_2 + H^+$

In *p*-chlorobenzoic acid, the inductive effect of chlorine in the para position is likely to cause electron deficiency on the α-cabon. The result of this would be that decarboxylation would not occur in acid medium and would be facilitated in neutral and basic media. This would be in harmony with the view of Fairclough[14], according to whom in trichloroacetic and tribromoacetic acids the α-carbon is attached to strongly electron attracting chloro and bromo groups which weakens the carbon-carboxy bond and facilitates its unimolecular rupture:



According to Fieser and Fieser[16], the inductive effect estimated from dipole moments accounts well for the activating or deactivating influence of different groups, but does not explain the orientation and it is seen that although the majority of electron attracting groups orient an entering group to the meta position, halogen substituents, like the electron repelling amino and hydroxyl groups are ortho, para directing. This is due to the resonance effect. Thus in chlorobenzene resonance leads to development of electron density at the ortho and para positions (with reference to chlorine atom). In *p*-chlorobenzoic acid, structures in which carboxyl also partakes of the resonance are possible. But according to Schubert[4], in the presence of substituents, the resonance interaction of the carboxyl with the benzene nucleus is inhibited. This leads to the weakening of the bond between the carboxyl and α-carbon and to its easy rupture. Another interesting explanation is that of Royals[15], according to whom the presence of meta orienting groups like carboxyl produces an electromeric effect indicated in the accompanying structure, due to which nucleophilic reactivity is developed at the α-carbon atom adjacent to meta orienting group. Hence an electrophilic reagent attacks this position and displaces the meta orienting group:



This also accounts satisfactorily the decarboxylation of benzoic acid[8] in the solvent resorcinol.

The bimolecular electrophilic substitution (S_E2) mechanism also accounts for the behavior of *p*-chlorobenzoic acid in the solvent glycerol which has been recently used by Clark[16] in decarboxylations. Glycerol being a polyhydric compound may act as a protophilic solvent and facilitate the ionization of *p*-chlorobenzoic acid to a limited extent.

$$\underset{\text{acid}}{HA} + \underset{\text{solvent}}{ROH} \rightleftharpoons (ROH_2)^+ + A^-$$

where $(ROH_2)^+$ is solvated hydrogen ion. The "hydrogen ions" so formed will facilitate the decarboxylation:

$$C_6H_4(Cl)(COOH) + H^+ \rightarrow C_6H_5Cl + CO_2 + H^+$$

The trends in the rate constants during decarboxylation in glycerol may be due to the fact that along with decarboxylation, the ester formation between glycerol and the acid takes place and the ester so formed decomposes decarboxylatively introducing complications in the interpretation of rate data. The retarding effect of aniline on decarboxylation in glycerol is also consistent with the S_E2 mechanism proposed, since aniline removes the hydrogen ions present and thus causes retardation in their attack on the carboxylic acid. On the contrary if decarboxylation were to occur via the first order decomposition of the carboxylate anion by the S_E1 mechanism, the rate should have increased due to the carboxylate anions being formed in abundance:

$$C_6H_5NH_2 + HA = (C_6H_5NH_3)^+ + A^-$$

Summary

The decarboxylation of *p*-chlorobenzoic acid has been studied in the solvent resorcinol. The activation energy, frequency factor and free energy and entropy of activation have been calculated. It is shown that the reaction follows the S_E2 mechanism.

Department of Chemistry
University College of Science
Osmania University
Hyderabad, India

14) R. A. Fairclough, *J. Chem. Soc.*, **1938**, 1186.
15) E. Royals, "Advanced Organic Chemistry" Constable & Co., p. 450.
16) L. W. Clark, *J. Am. Chem. Soc.*, **77**, 3130 (1955).

Sur la Dissolution des Minerais Sulfurés en Divers Milieux. III. Facteurs Intervenant à la Dissolution de la Chalcopyrite

Par Masami ICHIKUNI

(Reçu le 1er Novembre 1959)

On a examiné dans le rapport précédent[1] les influences du pH et de l'oxygène sur la dissolution de la chalcopyrite qui existe indépendamment des autres minéraux. Cependant tous les minerais produits dans la nature sont d'un système complexe de plusieurs minéraux. On ne peut pas supposer que chaque minéral se comporte indépendamment des autres au cours de la dissolution du minerai; il doit y avoir évidemment les actions réciproques entre les minéraux ainsi que celles des ions en solution. Ohashi[2] a déjà traité l'aspect général de ce problème; toutefois, à ce sujet, notre connaissance reste encore peu satisfaisante.

Expérience

On examine ici les effets de la présence de la pyrite, qui se trouve associée très souvent à la chalcopyrite, et ceux des ions de fer divalent, qui se forment par la désagrégation. Le dispositif destiné à l'expérience est le même que précédemment décrit[1].

Effets de la Présence de la Pyrite.—On mélange bien la pyrite et la chalcopyrite, en proportion connue; les deux ont été utilisées dans les expériences du rapport précédent. De ce mélange bien pulvérisé, on met 1 g. dans 200 ml. d'acide chlorhydrique 0.01 N et traite pendant deux heures à la température de 80°C. Le prélèvement de 10 ml. de liqueur s'effectue avec l'interval d'une demi-heure après l'addition du mélange. L'air est introduit dans un ballon contenant de la liqueur et de la poudre des minéraux avec un débit de 1,5 l./min. Quant aux conditions d'expériences, voir le rapport précédent.

Effets des Ions de Fer Divalent en Solution.—On dissout une quantité connue de sulfate de fer(II), $FeSO_4 \cdot 7H_2O$, dans l'acide chlorhydrique 0.01 N, prend 200 ml. de cette solution, y ajoute 1 g. de chalcopyrite pulvérisée, et traite pendant deux heures. Le prélèvement de liqueur s'effectue de la même façon que l'expérience décrite ci-dessus. Dans la liqueur on introduit de l'air ou de l'azote avec un débit de 1,5 l./min., l'introduction de ce dernier ayant pour but de réaliser une condition non-oxydante.

Résultats et Discussion

Co-présence de la Pyrite.—On a préparé quatre mélanges de proportions différentes de ces deux minéraux. La composition des mélanges est donnée dans le Tableau I. Les concentrations de fer divalent, de cuivre et de sulfate, qu'on a calculées en tenant compte de la diminution du volume de liqueur par l'évaporation, sont portées sur le Tableau II.

TABLEAU I. COMPOSITION DES MÉLANGES CONSISTANT EN PYRITE ET EN CHALCOPYRITE

No. du mélange	Pyrite	Chalcopyrite	Cu	Fe	$(Cu/Fe)_{min.}$
	%	%	%	%	$\times 10^{-2}$
1	95,5	4,5	1,56	45,8	3,71
2	84,6	15,4	5,34	44,0	12,1
3	62,1	37,9	13,1	40,4	32,4
4	51,4	48,6	16,8	38,6	43,6

Remarque: Les teneurs en cuivre et en fer des mélanges sont calculées d'après les résultats d'analyse des minéraux. La valeur $(Cu/Fe)_{min.}$ représente le rapport de teneur en cuivre sur celle en fer du mélange.

TABLEAU II. VARIATION DE CONCENTRATIONS DE FER DIVALENT, DE CUIVRE ET DE SULFATE AU COURS DE LA DISSOLUTION DE 1 g. DE CHAQUE MÉLANGE

No. du mélange	Durée du traitement hr.	Cu mg./l.	Fe^{2+} mg./l.	SO_4^{2-} mg./l.
1	0,5	3,92	16,3	29
	1	6,05	19,2	39
	1,5	9,89	23,6	49
	2	12,0	27,1	59
2	0,5	4,39	18,4	23
	1	6,89	22,5	30
	1,5	10,5	26,4	42
	2	14,4	31,0	53
3	0,5	4,13	19,4	17
	1	6,06	25,8	25
	1,5	8,12	31,8	32
	2	8,85	38,1	41
4	0,5	3,38	17,5	19,5
	1	4,00	23,6	24
	1,5	4,95	31,0	28
	2	5,95	37,6	36

Avec les données du Tableau II, on calcule la vitesse de dissolution à partir de la pente de la droite représentant la concentration de composants en fonction de durée du traitement. Sa vitesse est considérablement grande au

1) M. Ichikuni, Ce Bulletin, 33, 1052 (1960).
2) S. Ohashi, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), 74, 845 (1953).

Tableau III. Vitesse de dissolution du cuivre, du fer divalent et du sulfate à partir de 1 g. de mélange

No. du mélange	Cu mg./hr.	Fe^{2+} mg./hr.	SO_4^{2-} mg./hr	$(Cu/Fe)_{soln.}$ $(\times 10^{-2})$	$\frac{(Cu/Fe)_{soln.}}{(Cu/Fe)_{min.}}$
1	1,10	1,47	4,0	74,9	20,2
2	1,35	1,66	4,05	81,4	6,73
3	0,76	2,41	3,2	31,5	0,973
4	0,41	2,70	2,2	15,2	0,349

début, et après elle prend une valeur constante. Cette irrégularité de dissolution fait supposer la présence des parties facilement solubles du minéral, mais dont l'essentiel reste encore peu clair.

En examinant la vitesse de dissolution des composants sur les quatre échantillons de mélanges, on a trouvé des faits suivants: avec l'augmentation de proportion de chalcopyrite dans les mélanges,

1° la vitesse de dissolution du fer augmente,

2° celle du cuivre va d'abord montant, atteint le maximum, et après, descend graduellement.

3° celle du sulfate ne fait que diminuer.

Cela signifie que la pyrite et la chalcopyrite ne se dissolvent pas indépendamment l'une de l'autre et qu'il existe les influences réciproques sur la dissolution de leur mélange. D'où vient que la vitesse de dissolution des composants à partir du mélange est toujours différente de celle prévue d'après les résultats de la dissolution de chacun des deux minéraux purs.

La vitesse de dissolution des composants est montrée dans le Tableau III, où $(Cu/Fe)_{soln.}$ et $(Cu/Fe)_{min.}$ représentent respectivement le rapport de vitesse de dissolution du cuivre sur celle du fer, et le rapport de teneur en cuivre sur celle en fer du mélange.

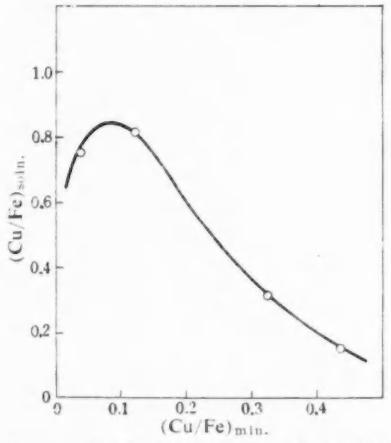


Fig. 1. Relation entre la valeur de $(Cu/Fe)_{min.}$ et la valeur de $(Cu/Fe)_{soln.}$

L'augmentation de teneur en cuivre du mélange, c'est-à-dire celle de valeur de $(Cu/Fe)_{min.}$, diminue la valeur de $(Cu/Fe)_{soln.}$. La Fig. 1 montre cette relation où la valeur de $(Cu/Fe)_{soln.}$ atteint le maximum au moment où la valeur de $(Cu/Fe)_{min.}$ est de 0,1 environ. La diminution de la valeur de $(Cu/Fe)_{min.}$ au-dessous de 0,1 abaisse également la valeur de $(Cu/Fe)_{soln.}$; mais le rapport de $(Cu/Fe)_{soln.}$ sur $(Cu/Fe)_{min.}$ a une signification particulière au point de vue géochimique et pratique.

Ce rapport soit r,

$$r = \frac{(Cu/Fe)_{soln.}}{(Cu/Fe)_{min.}}$$

Si $r > 1$, le minerai résiduel s'appauvrit en cuivre. Si $r < 1$, le minerai résiduel devient riche en cuivre. On a déjà remarqué cette possibilité d'enrichissement du cuivre dans certaines discussions concernant la dissolution de la chalcopyrite. Le point où $r = 1$ se trouve à la valeur de $(Cu/Fe)_{min.}$ de 0,32. Supposons que le minerai consiste simplement en pyrite et en chalcopyrite, alors la teneur en cuivre du minerai est de 13%. Comme la plupart des minerais sulfurés cuprifères sont moins riches en cuivre que ce minerai en supposition, l'enrichissement en cuivre du minerai résiduel n'aura pas lieu en réalité au cours de la désagrégation, et l'extraction du cuivre s'effectue efficacement par l'action des eaux.

Le rapport r représente l'efficacité d'extraction du cuivre à partir du minerai, ou le coefficient de concentration du cuivre de la phase solide en phase liquide; la valeur r est d'autant plus grande que la teneur en cuivre du minerai est plus petite (voir le Tableau III). Il est donc avantageux d'extraire du cuivre d'un minerai pauvre en cuivre par l'action des eaux. Ce procédé est exécuté dans certain nombre de mines. A titre d'exemple, on calcule ici la valeur r sur le minerai siliceux (kéi-kô), qui a été utilisé comme échantillon dans le premier rapport relatif à la dissolution des minerais sulfurés[3], où

$$(Cu/Fe)_{min.} = 0{,}361 \times 10^{-2}$$

$$(Cu/Fe)_{soln.} = 9{,}83 \times 10^{-2}$$

3) M. Ichikuni et H. Kamiya, Ce Bulletin, 32, 1368 (1959).

on a donc $r=27,2$, qui est exceptionnellement grand ; on notera d'ailleurs que la teneur en cuivre du minerai n'est que de 0,048%.

Effets des Ions de Fer Divalent en Solution. —Le Tableau IV donne les concentrations de composants dissous, correspondant à l'heure du prélèvement de liqueur et à la quantité additionnée de fer divalent. La quantité dissoute de fer et de sulfate est très limitée et reste inférieure à l'erreur du dosage de ces composants, lorsque la concentration de fer est plus de 500 mg./l. Notons que le fer est ajouté sous forme du sulfate de fer(II), $FeSO_4 \cdot 7H_2O$. Cela gêne de vérifier l'augmentation de concentration de sulfate.

Le dosage du fer divalent et du sulfate s'est donc effectué, lorsque la concentration initiale de fer était inférieure à 100 mg./l. On montre dans le Tableau V la vitesse de dissolution du fer divalent, du cuivre et du sulfate calculée d'après le Tableau IV.

TABLEAU IV. VARIATION DE CONCENTRATIONS DE CUIVRE, DE FER DIVALENT ET DE SULFATE AU COURS DE LA DISSOLUTION DE 1 g. DE CHALCOPYRITE, EN PRÉSENCE DU FER DIVALENT

Conc. initiale de Fe^{2+} mg./l.	Durée du traitement hr.	Cu mg./l.	Fe^{2+} mg./l.	SO_4^{2-} mg./l.
30	0,5	0,16	32,0	37
	1	0,45	35,0	37
	1,5	0,45	38,5	37
	2	0,43	40,5	38
55	0,5	<0,10	57,5	78
	1	0,15	61,2	80
	1,5	0,59	63,6	82
	2	0,40	67,2	84
100	0,5	0,15	102	150
	1	0,32	104	150
	1,5	0,69	106	155
	2	0,95	112	157
500	0,5	3,74	—	—
	1	8,10	—	—
	1,5	11,6	—	—
	2	15,3	—	—
1000	0,5	5.81	—	—
	1	10,6	—	—
	1,5	16,6	—	—
	2	23,2	—	—
1000	0,5	0,34	—	—
	1	<0,10	—	—
	1,5	<0,10	—	—
	2	<0,10	—	—

Remarque : Ces expériences ont été effectuées en présence de l'air, à l'exception de la dernière qui a été exécutée en milieu exempt d'oxygène.

TABLEAU V. VITESSE DE DISSOLUTION DU CUIVRE, DU FER DIVALENT ET DU SULFATE À PARTIR DE 1 g. DE CHALCOPYRITE EN PRÉSENCE DU FER DIVALENT

Conc. initiale de Fe^{2+} mg./l.	Cu mg./hr.	Fe^{2+} mg./hr.	SO_4^{2-} mg./hr.
30	<0,05	1,05	0,15
55	<0,05	1,28	0,80
100	0,11	1,33	0,90
500	1,55	—	—
1000	2,28	—	—
1000	0	—	—

Remarque : Le dernier rang correspond à l'expérience effectuée en milieu exempt d'oxygène.

Dissolution du Fer.—A mesure que la concentration de fer divalent augmente, la quantité dissoute de fer à partir de la chalcopyrite augmente. Cette tendance existerait même en présence du fer dont la concentration est plus de 100 mg./l. D'autre part, sous les conditions où la concentration de fer est moins de 100 mg./l., il n'y a presque pas de dissolution du cuivre. On peut donc admettre que la dissolution du fer s'effectue plus facilement que celle du cuivre. L'accélération de dissolution du fer serait attribuable à des ions de fer trivalent. On a donc :

$$(1)\quad CuFeS_2 + 2Fe^{3+} \rightarrow CuS + S + 3Fe^{2+}$$

et on en déduit, en tenant compte de l'augmentation de concentration de sulfate :

$$(2)\quad CuFeS_2 + 8Fe^{3+} + 4H_2O \rightarrow CuS + SO_4^{2-} + 8H^+ + 9Fe^{2+}$$

Toutefois l'augmentation de concentration de sulfate est toujours inférieure à celle de fer,

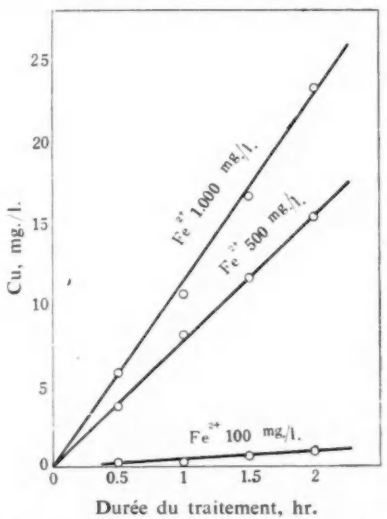


Fig. 2. Dissolution du cuivre à partir de la chalcopyrite en présence du fer divalent et de l'air.

par conséquent la réaction exprimée par l'équation (1) est prédominante dans le cas d'expériences.

Dissolution du Cuivre.—Elle n'est presque pas remarquée en présence du fer dont la concentration est inférieure à 100 mg./l. Au delà de 100 mg./l., pourtant, le cuivre se dissout considérablement. La Fig. 2 est dessinée avec les chiffres du Tableau IV. On y voit bien la dissolution du cuivre en présence du fer dont la concentration est supérieure à 100 mg./l. D'après Ohashi[2], la dissolution du cuivre à partir de la chalcopyrite est favorisée par l'action des ions de fer trivalent, ce qui est à comparer avec le résultat décrit ci-dessus. Un mécanisme de dissolution du cuivre, semblable à celui du fer, peut être admis:

$$(3)\quad CuFeS_2 + 4Fe^{3+} \rightarrow Cu^{2+} + 5Fe^{2+} + 2S$$

La quantité de fer trivalent est d'autant plus grande que celle de fer divalent en solution est plus grande, car le fer trivalent est formé par l'oxydation du fer divalent par l'oxygène. Ce phénomène favorise la dissolution de la chalcopyrite dans une mesure considérable.

Ce résultat est soutenu par le fait que la présence d'une quantité importante de fer divalent n'accélère pas la dissolution du cuivre en milieu exempt d'oxygène.

La majeure partie de fer en solution se trouve à l'état divalent, et la quantité de fer trivalent n'est au maximum que de 1% de celle de fer divalent, si la solution est en contact avec un minerai. Cela s'explique bien par le fait que la réaction de l'attaque du fer trivalent sur le minéral s'effectue rapide par rapport à celle de l'oxydation du fer divalent en fer trivalent.

On n'a pas de chiffres en ce qui concerne le soufre, mais il est probable que le fer trivalent intervient à la formation du sulfate. Comme les eaux en contact avec le gisement des minerais sulfurés contiennent généralement une quantité considérable de fer, la dissolution de la chalcopyrite ne peut pas s'expliquer par un mécanisme simple, mais elle est soumise à des influences de plusieurs facteurs, dont dépend la quantité dissoute de cuivre, de fer et de soufre.

Résumé

La dissolution du cuivre à partir de la chalcopyrite est influencée par le fer en solution qui est en contact avec le minéral, mais plus remarquable est l'effet de la co-présence de la pyrite. L'influence du fer en solution devient considérable, si la concentration de fer dépasse 100 mg./l.

L'auteur exprime ses remerciements profonds à Monsieur le professeur Noguchi qui a bien voulu lire ce rapport et lui donner des conseils utiles, et particulièrement à Monsieur H. Kamiya qui a participé aux expériences et à la discussion sur l'interprétation des résultats.

Faculté des Sciences
Université Métropolitaine de Tokyo
Sétagaya, Tokyo

Sur la Dissolution des Minerais Sulfurés en Divers Milieux. IV. Dissolution du Minerai Jaune

Par Masami ICHIKUNI

(Reçu le 19 Novembre 1959)

La dissolution des minerais sulfurés, notamment celle de la pyrite et de la chalcopyrite, a été systématiquement étudiée, et le résultat des expériences a été réuni dans le 2ème rapport[1] des travaux de la présente série. L'action réciproque des deux minéraux lors de la dissolution de leur mélange était le sujet du rapport précédent[2]. Il serait donc intéressant de vérifier que la dissolution d'un certain minerai peut être expliquée complètement par l'application des résultats obtenus jusqu'ici. Comme les minerais sont d'une association de nombreux minéraux, leur dissolution est influencée à la fois par les facteurs de milieu et par l'action réciproque existant entre les minéraux. Il n'y aurait pas de manière générale d'interpréter le résultat d'expériences concernant la dissolution des minerais. L'état actuel des

1) M. Ichikuni, Ce Bulletin, 33, 1052 (1960).
2) M. Ichikuni, Ce Bulletin, 33, 1159 (1960).

recherches ne permet que d'appliquer les résultats obtenus sur les minéraux ou leur mélange à l'interprétation de la dissolution d'un minerai sulfuré qui a une composition minéralogique simple. On prend donc comme échantillon le minerai jaune, ou *ô-kó*, qui est constitué généralement par le quartz, la pyrite et la chalcopyrite, sans parler des composants mineurs.

Expérience

On suit la méthode d'expériences qui est déjà mise au point et dont la description détaillée est donnée dans le 2ème rapport.

L'échantillon utilisé dans cette étude est le minerai jaune qui s'est produit à la mine de Kosaka, département d'Akita. Le résultat d'analyse chimique en est montré par le Tableau I. On a décelé par voie spectrographique des traces d'aluminium, d'argent et d'arsenic. Les minéraux arsénifères existent en quantité négligeable. On a soumis ensuite le minerai à l'examen diffractométrique à rayons X, et a décelé le quartz, la pyrite et la chalcopyrite. La présence de la blende est incertaine, car la teneur en zinc est trop petite pour qu'on identifie précisément l'espèce de minéral de zinc. Toutefois il est probable que le zinc se trouve sous forme de blende. En admettant que

Tableau I. Composition chimique du minerai jaune

Composant	Teneur %
Fe	31,17
Cu	2,41
Zn	0,14
S	35,25
SiO_2	29,13
Total	98,10

Tableau II. Composition minéralogique du minerai jaune

Minéral	Formule chimique	Teneur %
Pyrite	FeS_2	62,4
Chalcopyrite	$CuFeS_2$	7,0
Blende	ZnS	0,2
Quartz	SiO_2	29,1
Total		98,7

chaque minéral a une composition correspondant à la formule chimique idéale, on a calculé la composition minéralogique du minerai en question, dont le résultat est porté sur le Tableau II. La présence d'une seule espèce de minéral de cuivre simplifie l'analyse des résultats d'expériences.

Pour chaque expérience, on a utilisé 1 g. de minerai pulvérisé.

Résultats et Discussion

Expérience de Dissolution Effectuée en Présence de l'Air.—Le Tableau III rassemble les résultats obtenus. La vitesse de dissolution du fer, de cuivre, du zinc et de soufre a été calculée en traçant la droite de concentration de chaque composant en fonction de la durée du traitement. On envisage ici le soufre qui se dissout sous forme de sulfate. Le résultat du calcul est montré dans le Tableau IV. A mesure que le pH du milieu diminue, la vitesse de dissolution du cuivre et du zinc augmente, tandis que celle du sulfate ne fait que diminuer. La vitesse de dissolution du fer est approximativement constante en milieu à pH inférieur à 3. A partir de ces résultats, on conclut que la dissolution du minerai jaune ressemble à celle de la pyrite que l'on a étudiée sous la même condition. On notera que ce

Tableau III. Variation de concentration de fer divalent, de cuivre, de zinc et de sulfate au cours de la dissolution du minerai jaune en présence de l'air

Milieu	Valeur moyenne de pH	Durée du traitement hr.	Fe^{2+} mg./l.	Cu mg./l.	Zn mg./l.	SO_4^{2-} mg./l.
HCl 0,1 N	1,1	1	62,5	31,4	4,6	111
		2	66,8	34,3	5,9	111
		3	68,0	35,7	6,2	114
		4	70,0	36,4	6,1	113
HCl 0,01 N	2,05	1	37,7	14,2	3,0	97
		2	40,7	15,2	3,5	104
		3	41,8	16,2	3,5	110
		4	42,9	17,0	3,6	116
HCl 0,001 N	3,0	1	36,7	9,7	3,2	87,5
		2	38,1	9,8	3,3	98
		3	42,0	9,8	3,3	116
		4	43,7	9,7	3,5	133
H_2O	3,0*	1	16,7	5,4	2,8	77
		2	16,5	6,2	2,3	97
		3	16,6	6,6	4,5?	113
		4	16,7	6,8	2,8	169

* L'abaissement du pH est dû à la formation de l'acide sulfurique par l'oxydation.

Tableau IV. Vitesse de dissolution du fer divalent, du cuivre, du zinc et du sulfate à partir de 1 g. de minerai jaune en présence de l'air

Milieu	Fe^{2+} mg./hr.	Cu mg./hr.	Zn mg./hr.	SO_4^{2-} mg./hr.	Cu/Fe^{2+}	Zn/Fe^{2+}
HCl 0,1 N	0,34	0,24	>0,26	0,14	0,71	>0,89
HCl 0,01 N	0,39	0,20	>0,10	1,24	0,51	>0,26
HCl 0,001 N	0,44	0	0,03	3,0	0	0,07
H_2O	0	0	0	3,6	—	—

minéral constitue la plus grande partie (62,4%) du minerai.

Fer.—Lors du traitement du minerai avec de l'eau, la dissolution du fer n'a pas lieu, bien que le pH de la liqueur s'abaisse à 3 près. Cependant sous la même condition, une quantité appréciable de fer s'est dissoute à partir de la pyrite. On envisage maintenant le problème du comportement de la pyrite qui se trouve associée à d'autres minéraux. Ce problème pourrait être résolu par l'examen des influences des minéraux co-existants. D'ailleurs il est bien connu que la présence de la blende empêche la dissolution de la pyrite et de la chalcopyrite[2]. Pourtant cette raison n'est pas très convaincante pour la résolution totale du problème.

Cuivre.—La dissolution du cuivre est évidemment influencée par le pH du milieu, mais d'une façon différente du cas de dissolution du cuivre à partir de la chalcopyrite qui se trouve indépendamment des autres minéraux. Dans le cas de la chalcopyrite, la vitesse de dissolution du cuivre ainsi que du fer augmente brusquement au moment où l'acidité du milieu atteint 0,1 N, tandis que celle du cuivre à partir du minerai jaune n'augmente que légèrement par l'abaissement du pH de 2 à 1. La dissolution sélective de la chalcopyrite ne se produit pas même à pH 1. La valeur de $(Cu/Fe)_{soln.}$, rapport de vitesse de dissolution du cuivre sur celle du fer, était de 0,38 à pH 1 pour la chalcopyrite. Elle est de 0,71 au même pH pour le minerai, et dépasse le chiffre trouvé pour la chalcopyrite.

On constate ainsi que la dissolution du cuivre à partir de la chalcopyrite est certainement favorisée par la présence de la pyrite, ce qui conforme au résultat du rapport précédent[2]. Les effets de la pyrite sur la dissolution de la chalcopyrite qui ont été trouvés en milieu à pH 2 continuent à exister même à pH 1. Le coefficient de concentration du cuivre par rapport au fer, au cours de la dissolution du minerai, est exprimé par la relation suivante et calculé d'après le résultat d'expériences effectuées en milieu acide 0,01 N. Le coefficient soit r, et on a:

$$r=\frac{(Cu/Fe)_{soln.}}{(Cu/Fe)_{min.}}=\frac{51\times10^{-2}}{7{,}75\times10^{-2}}=6{,}6$$

où $(Cu/Fe)_{soln.}$ est le rapport de vitesse de dissolution du cuivre sur celle du fer, et $(Cu/Fe)_{min.}$ est le rapport de teneur en cuivre du minerai sur celle en fer. Ce chiffre n'est pas très éloigné de 10,6, chiffre prévu d'après les expériences effectuées sur le mélange de la pyrite et de la chalcopyrite. Toutefois il n'est pas facile à dire par quelle raison cet écart peut être expliqué. Comme cet écart peut être attribué, d'une part, aux effets des substances co-existantes, telles que la blende, et d'autre part, à l'individualité du minerai, une étude plus approfondie est exigée pour la mise en évidence de la cause de l'écart. Il est probable que les impuretés contenues dans le minerai détermine son individualité, par conséquent on ne peut pas passer des études des composants mineurs.

D'après le Tableau IV, la valeur de $(Cu/Fe)_{soln.}$ est d'autant plus grande que le pH du milieu est plus petit, d'où on désigne comme facteurs intervenant à la dissolution du cuivre à partir de la chalcopyrite située dans la zone d'oxydation les suivants: pH, co-présence de la pyrite et concentration de fer en solution. Comme il a été signalé dans le rapport précédent,[2] l'influence de fer en solution devient remarquable, lorsque sa concentration dépasse 100 mg./l. Il faut qu'on tienne compte de la température, de la pression et de la co-présence des autres substances que la pyrite, mais l'état actuel des recherches ne permet pas de discuter davantage sur ces facteurs de milieu.

Zinc.—Le travail effectué par Ohashi a montré que la dissolution de la blende est accélérée par la présence de la pyrite et de la chalcopyrite[3]. Les effets du pH se joignant à ceux de la présence de ces minéraux, le phénomène de la dissolution du zinc a un aspect très compliqué, par conséquent la formation de l'acide par la désagrégation des minerais sulfurés est un facteur non-négligeable, tant qu'on traite la dissolution de ce métal.

Expérience de Dissolution Effectuée en Milieu Exempt d'Oxygène.—Le résultat obtenu est porté sur le Tableau V. Comme on a prévu d'après les résultats d'expériences concernant

3) S. Ohashi, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **74**, 845 (1953).

TABLEAU V. VARIATION DE CONCENTRATION DE FER DIVALENT, DE CUIVRE, DE ZINC ET DE SULFATE AU COURS DE LA DISSOLUTION DU MINERAI JAUNE EN MILIEU EXEMPT D'OXYGÈNE

Milieu	Valeur moyenne de pH	Durée du traitement hr.	Fe^{2+} mg./l.	Cu mg./l.	Zn mg./l.	SO_4^{2-} mg./l.
HCl 0,1 N	1,2	1	37,1	6,75	2,9	74
		2	36,7	6,30	3,4?	76
		3	36,9	5,10	2,8	76
		4	36,3	4,58	2,6	74
HCl 0,01 N	2,05	1	37,8	2,25	2,2	77
		2	39,7	0,90	2,0	78
		3	40,1	0,44	2,0	81
		4	40,0	0,36	1,6	81
HCl 0,001 N	3,0	1	37,0	0,61	2,6	78
		2	37,8	0,59	2,5	79
		3	38,1	0,50	2,4	79
		4	38,3	0,26	2,2	81
H_2O	4,0	1	32,4	0,2	2,4	81
		2	33,2	0,2	1,9	82
		3	33,0	0,2	1,7	83
		4	33,6	0,2	1,6	84

la dissolution de la pyrite et de la chalcopyrite, la formation du sulfate n'a pas lieu. Au reste, l'indissolution du fer attire notre attention. Il peut y avoir là les influences de la co-présence des minéraux. On remarque encore que sous les conditions acides et non-oxydantes le comportement du cuivre est différent de celui du zinc.

Cuivre.—L'echantillon utilisé étant partiellement oxydé par la pulvérisation, une petite quantité de cuivre se dissout au début du traitement, même en milieu exempt d'oxygène. La concentration de cuivre diminue, pourtant, avec la durée du traitement. De ces expériences on ne peut tirer les matières suffisantes pour avoir des idées claires sur le mécanisme de la fixation du cuivre. Il se peut que l'hydrogène sulfuré, formé par l'action de l'acide sur le minerai sulfuré, cause la précipitation ou l'adsorption du cuivre sous forme de sulfure. Ce phénomène se produit même à pH 1, ce qui souligne que la dissolution du cuivre à partir des minerais sulfurés est peu possible dans les conditions acides et non-oxydantes, et que le contact des eaux contenant des ions de cuivre avec un minerai sulfuré conduit à la fixation du cuivre, si le milieu est caractérisé par l'absence de l'oxygène. On reconnaît la fixation de ce métal dans le phénomène de l'enrichissement secondaire du minerai cuprifère. Bien que cette fixation est réalisée par l'expérience, il est difficile d'éclaircir par le moyen de laboratoire les réactions qui prennent part à l'enrichissement secondaire du cuivre.

Zinc.—Le comportement du zinc en milieu à pH 1 n'a pas de ressemblance avec celui du cuivre: le zinc n'a pas de tendance à être fixé au cours du traitement. A pH supérieur à 2, la concentration de zinc commence à diminuer avec la durée du traitement. On en déduit la précipitation ou l'adsorption du zinc sous forme de sulfure, analogue à celle du cuivre. Toutefois la vitesse de fixation du zinc est moins grande que celle du cuivre. Ce fait est expliqué par la différence de la propriété chimique des deux métaux.

L'explication des réactions chimiques participant à l'enrichissement secondaire et l'examen des facteurs de milieu intervenant à ce phénomène composeront la suite des études de cette série.

Résumé

Les facteurs favorisant la dissolution du cuivre à partir de la chalcopyrite sont les suivants: condition acide et oxydante, co-présence de la pyrite et contact avec l'ion de fer. Ils favoriseraient également la dissolution du zinc à partir de la blende. On peut bien expliquer la dissolution du minerai jaune, en prenant ces facteurs en considération.

L'auteur se permet d'adresser ses remerciements sincères à Monsieur le professeur Noguchi qui a bien voulu lire le manuscrit de ce rapport et aux chercheurs du laboratoire qui lui ont conseillé au cours de l'exécution de ce travail. Il remercie particulièrement Monsieur H. Kamiya pour la collaboration précieuse. Il remercie aussi Monsieur le professeur T. Watanabé et Monsieur le docteur A. Kato, à l'Université de Tokyo, pour l'examen diffractométrique à rayons X de l'échantillon ainsi que les renseignements utiles sur les propriétés des minerais sulfurés.

Faculté des Sciences
Université Métropolitaine de Tokyo
Sétagaya, Tokyo

Synthesis of 2-Phenyltropone and its Bromo-derivatives from 2-Phenylsuberone[1]

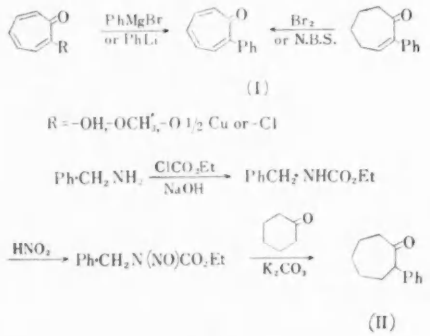
By Tadashi MUROI*

(Received December 22, 1959)

Nozoe, Mukai and their collaborators[2,3] have already reported that 2-phenyltropone (I) is obtained by the reaction of tropolone methyl ether and phenylmagnesium bromide in a good yield. Doering and others[4,5] have also obtained I by the treatment of tropolone and its derivatives with phenylmagnesium bromide or phenyllithium. On the other hand, it is well known that[6-9] various troponoid compounds are derived from cycloheptanons by their direct bromination-dehydrobromination. Nozoe and Ito[10,11] have obtained bromo-derivatives of I by the reaction of bromine and 2-phenylcyclohept-2-enone derived from cycloheptanone. Ginsburg and others[12] have also obtained I by the application of *N*-bromosuccinimide to 2-phenylcycloheptenone.

The present author has examined the reaction of 2-phenylsuberone (II) and bromine, and obtained I and its bromo-derivatives in a good yield. This paper describes these results.

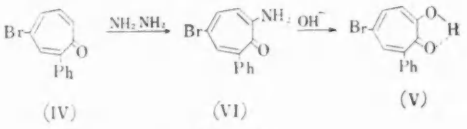
2-Phenylsuberone (II) was prepared by the treatment of cyclohexanone with phenyldiazomethane derived from benzylamine according to Gutsche's method[13].



Scheme 1

When three molar equivalents of bromine were added dropwise into a solution of II dissolved in glacial acetic acid, the reaction of bromine proceeded smoothly with the liberation of hydrogen bromide, and yellow needles (III), decomp. p. 185~187°C, were yielded. By the treatment with water, III was decomposed to give yellow needles (IV) m. p. 91~92°C. The yield of IV is about 30% calculated from II. From the mother liquor, a small amount of I was obtained.

Previously, Nozoe and others[10] confirmed the structure of 4-bromo-2-phenyltropone (IV) by deriving from it 5-bromo-3-phenyltropolone (V) through 7-amino-4-bromo-2-phenyltropone (VI). The present author has also obtained V from IV by the same treatment, and the melting point and the ultraviolet spectrum of IV (Fig. 1) are also identical with those of 4-bromo-2-phenyltropone[10]. It is obvious, therefore, that III is a hydrogen bromide adduct of 4-bromo-2-phenyltropone.



Scheme 2

In some cases, especially, when the reaction was carried out at a lower temperature, a small amount of colorless needles (VII), m. p. 155~156°C and colorless needles (VIII), m. p. 128~129°C were obtained besides IV and I. Though the details of the structure of VII and VIII are not clarified yet, each of them has been proved to be a dibromo-derivative of II from their analytical value.

When four molar equivalents of bromine

1) Presented at the Local Meeting of Tohoku District of the Chemical Society of Japan, held at Kōriyama, Oct. 5. 1957.

* Present address, Hitachi Research Laboratory, Hitachi, Ltd. Sukekawa-machi, Hitachi-shi.

2) T. Nozoe, T. Mukai and J. Minegishi, *Proc. Japan Acad.*, **27**, 419 (1951).

3) T. Nozoe, T. Mukai, J. Minegishi and T. Fujisawa, Science Repts., *Tohoku Univ.*, **I**, 388 (1953).

4) W. von E. Doering and C. H. Hyskey, *J. Am. Chem. Soc.*, **74**, 5688 (1952).

5) W. von E. Doering and J. R. Mayer, ibid., **75**, 2387 (1953).

6) H. J. Dauben and H. J. Ringold, ibid., **73**, 876 (1951).

7) T. Nozoe, Y. Kitahara, T. Ando and S. Masamune, *Proc. Japan Acad.*, **27**. 415 (1951).

8) T. Nozoe, Y. Kitahara, T. Ando, S. Masamune and H. Abe, *Science Repts. Tohoku Univ.*, **I**, **36**, 166 (1952).

9) P. L. Pauson, *Chem. Revs.*, **55**, 9, (1955).

10) T. Nozoe, S. Ito and K. Sonobe, *Science Repts, Tohoku Univ.*, **I**, **38**, 141 (1954).

11) T. Nozoe, S. Ito and K. Sonobe, *Proc. Japan Acad.*, **29**, 101 (1953).

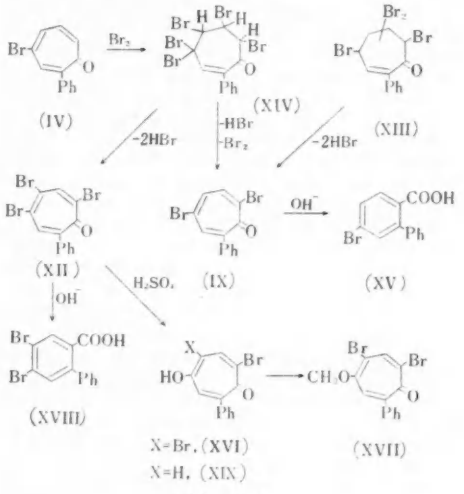
12) D. Elead and D. Ginsburg, *J. Chem. Soc.*, **1954**, 471.

13) D. Gutsche, "Organic Syntheses", **35**, 91 (1955).

were applied to II, it was found that the yield of IV increased remarkablly to about 80%. In this case, moreover, yellow needles (IX), m. p. 117~118°C, and yellow acid crystals (X), decomp. p. ca 200°C, were obtained, though in a small amount.

The reaction of five molar equivalents of bromine and II was found to be somewhat complicated. In this case, prolonged heating and stirring were required to complete the reaction, and the yield of IX was increased to about 78%. From the neutral portion of the mother liquor, besides IX and IV, the following compounds were obtained: pale yellow plates (XI) m. p. 79~80°C pale yellow needles (XII) m. p. 112~113°C and colorless needles (XIII) m. p. 149~150°C. From the acidic portion, a fair amount of X was obtained.

Previously, Mukai[14] reported that a dibromo-derivative of I was obtained by the decomposition of 4, 4, 5, 6, 7-pentabromo-2-phenyl-2-cyclohepten-1-one (XIV) which was derived from IV, and that its structure was assumed to be 4, 7-dibromo-2-phenyltropone.



Scheme 3

The author also obtained IX by the bromination of IV as Mukai reported, and moreover found that, when IX was treated with dilute alkali, it easily rearranged itself to give 4-bromo-2-phenylbenzoic acid (XV)[15]. Therefore, IX seems to be 4, 7-dibromo-2-phenyltropone as Mukai assumed.

The crystal (XI) was proved to be identical with the authentic sample of 7-bromo-2-phenyltropone[12] by the mixed melting point.

The crystal (XII) was found to be a tribromo-derivative of I from its analytical value and U. V. spectrum, (Fig. 1) and it was also obtained from XIV by treating with pyridine. When XII was heated for a long period with 75% sulfuric acid, it gave 5, 7-dibromo-4-hydroxy-2-phenyltropone (XVI)[16], whose structure was confirmed by deriving from it 5, 7-dibromo-4-methoxy-2-phenyltropone (XVII)[16]. It was also found that, by heating with dilute alkali, XII was easily rearranged to give colorless needles of m. p. 180~181°C of 4, 5-dibromo-2-phenylbenzoic acid (XVIII). These facts, therefore, indicated that XII was 4, 5, 7-tribromo-2-phenyltropone.

The crystal (XIII) was a tetrabromo-derivative of 2-phenylsuberenone. When XIII was heated to effect dehydrobromination, it gave IX. It is, therefore, found that two of four bromine atoms of XIII, at least, are presented in the C-4 and C-7 position, and XIII is probably one of the intermediate products in the formation of IX from II.

Though the detailed studies on the acidic product (X) will be described in the near future, it will be noted here that X is a mixture of 7-bromo-4-hydroxy-2-phenyltropone-(XIX) and XVI, and that they are isolated from each other with difficulty.

. Even when five molar equivalents of bromine were applied to II, the formation of X was very small, if the addition of bromine was carried out at 30~60°C and the duration of heating was shortened. In this case, the yield of IX was about 30% and that of XIII proved to be increased.

In order to prevent the formation of X completely, II was treated with five moles of bromine in a mixture of acetic acid and acetic anhydride but a considerable amount of X was still found in the acidic portion. In this case, IV was obtained mainly in about 30% yield but IX was in poor yield and, moreover, a trace of yellow needles (XX) m. p. 144~145°C was obtained. This compound XX was proved to be a monobromo-derivative of I from its analytical value and U. V. spectrum (Fig. 1) and gave by the further bromination a dibromo compound XXI as pale yellow needles m. p. 141~142°C.

As it is known that, on the bromination of tropone derivatives, the bromine atom enters at the carbon atoms adjacent to the carbonyl carbon (i. e., C-2 and C-7) preferentially[14,17], at least the one of bromine atoms of XXI was assumed to be in the C-7 position. It was proved, however, that XXI was identical

14) T. Mukai, *Science Repts. Tohoku Univ.*, **I**, **38**, 280 (1954).
15) I. M. Heilbron et al., *J. Chem. Soc.*, **1938**, 115.
16) T. Muroi, to be published.
17) T. Mukai, This Bulletin, **31**, 846 (1958).

with neither 3, 7-[18], 4, 7-, 5, 7-[14] and 6, 7-dibromo-derivative[14] of I. Therefore, the other bromine atom of XXI, i.e., that of XX is assumed not to be in the tropone nucleus but in the phenyl group.

In the above reactions, the yield of I was poor but, as it is known that IV is easily debrominated to give I, this reaction is considered to be a useful synthetic method for the preparation of I. As the mechanism of this reaction is not clarified yet this is considered to be an interesting question in future.

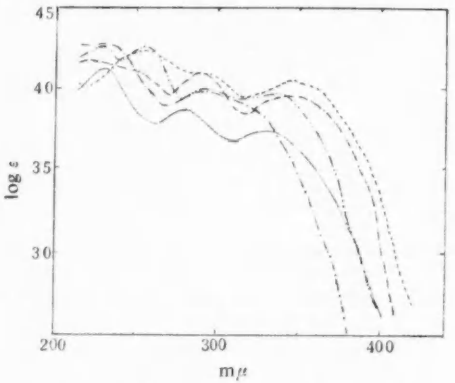


Fig. 1. U. V. Absorption spectra of IV, IX, XII, XX and XXI in methanol.
——— (IV) – – – (IX) ------ (XII)
–·–· (XX) –··–·· (XXI)

Experimental[19]

Action of Bromine on 2-Phenylsuberone (II).—II was prepared according to the description of "Organic Syntheses"[13], and the fraction of b. p. 110~111°C/0.3~0.4 mmHg; n_D^{17}: 1.5390, was used[20].

a) Application of 3 Molar Equivalents of Bromine.—i) To a solution of 4.7 g. of II in 7.5 cc. of acetic acid, a mixture of 12 g. of bromine and 2 cc. of acetic acid was added dropwise with stirring for a period of 1 hr. at 30~40°C. After the addition, the reaction mixture was heated on a water bath to about 80°C and kept at the same temperature for 30 min. and then yellow crystals were separated out. Additional heating and stirring were continued until there was no liberation of hydrogen bromide. After cooling, the crystals were collected by filtration and washed with a small amount of acetic acid and benzene, successively. Yellow crystals (III) decomp. p. 185~187°C, 3.7 g. were obtained. On standing overnight, the filtrate gave the second crop of III, 0.28 g. When III was treated with about 50 cc. of water, 2.3 g. of pale yellow crystals were obtained. Recrystallization from alcohol or methanol gave yellow needles (IV) m. p. 91~92°C.

Found: C, 60.06; H, 3.63. Calcd. for $C_{13}H_9OBr$: C, 59.74; H, 3.74%.

The mother liquor was evaporated under reduced pressure and the residue, treated with water, was extracted with chloroform. The extract was washed with 10% sodium carbonate solution. After the removal of the solvent, the residual oil, dissolved in a mixture of benzene and petroleum ether (1 : 1), was chromatographed through a silica-gel column. It was eluted successively with a mixture of benzene and petroleum ether, benzene, chloroform and alcohol. From the chloroform eluent, 0.2 g. of raw crystals of I was obtained but the other eluent gave no crystals.

Acidification of the sodium carbonate solution gave no crystalline products.

ii) As in the above case, 4.7 g. of II in 7.5 cc. of acetic acid was treated with 12 g. of bromine dissolved in 2 cc. of acetic acid. In this case, however, heating was stopped at ones when the separation of III occurred. After cooling, III was collected by filtration and treated with water to give 2.0 g. of IV. The filtrate, after standing for a few days, gave 0.22 g. of pale yellow crystals, m. p. 115~120°C. This was decomposed with water to give colorless crystals m. p. 175~180°C. Recrystallization from alcohol or benzene gave sparingly soluble colorless needles (VII) m. p. 155~156°C and comparatively soluble colorless needles (VIII) m. p. 128~129°C, respectively.

Found for VII: C, 45.41; H, 3.95. Calcd. for $C_{13}H_{14}OBr_2$: C, 45.12; H, 4.08%.

Found for VIII: C, 45.23; H, 3.97. Calcd. for $C_{13}H_{14}OBr_2$: C, 45.12; H, 4.08%.

After being treated as in i), the mother liquor gave 0.5 g. of I.

b) Application of 4 Molar Equivalents of Bromine.—To a solution of 4.7 g. of II in 7.5 cc. of acetic acid, 16 g. of bromine dissolved in 2.7 cc. of acetic acid was added dropwise during about 1.5 hr. at 30~60°C. In this case, when treated as in a), it gave three kinds of crystals from the neutral portion as follows: 5.2 g. of IV, 0.28 g. of IX and a small amount of I. Recrystallization of IX from alcohol gave yellow needles, m. p. 117~118°C.

Found: C, 45.68; 2.38. Calcd. for $C_{13}H_9OBr_2$: C, 45.93; H, 2.36%.

The acidic portion gave yellow crystals (X) decomp. p. 185~190°C, 0.02 g.

c) Application of 5 Molar Equivalents of Bromine.—i) To a solution of 4.7 g. of II in 7.5 cc. of acetic acid, 20 g. of bromine dissolved in 3.5 cc. of acetic acid was added dropwise over a period of 2.5 hr. at 30~65°C. After the addition, the mixture was heated on a water bath for 4 hr., until there was no liberation of hydrogen bromide. By chilling with ice no crystals were separated out, then the solvent was distilled off under reduced

18) T. Nozoe, T. Mukai and I. Murata, *Proc. Japan Acad.*, **28**. 142 (1952).

19) M. p. are uncorrected. The microanalyses and the measurement of U. V. spectra were carried out by Mrs. K. Kodaira of Tohoku University, to whom the author's deep gratitude is hereby expressed.

20) The fractional distillation of II was carried out at Kitahara's Laboratory in the Chemical Research Institute of Non-aqueous Solution, Tohoku University, using the Stedtmann type fractional column (40 plates). The author expresses hereby his deep gratitude to Professor Yoshio Kitahara for his kindness.

pressure. The residue, decomposed with water, was extracted with chloroform. After washing the extract with 10% sodium carbonate solution, the solvent was evaporated off. The residual oil (10 g.), digested with alcohol or ether, gave 0.9 g. of IX m. p. 113~115°C. After being treated as in a) i), the mother liquor was chromatographed through a silica-gel column. The benzene-petroleum ether eluent gave 1.0 g. of XIII, 0.8 g. of XII and 1.0 g. of IX, successively.

Recrystallization of XIII from alcohol gave colorless needles m. p. 149~150°C.

Found: C, 31.49; H, 2.13. Calcd. for $C_{13}H_{10}OBr_4$: C, 31.11; H, 2.00%.

Recrystallization of XII from alcohol gave pale yellow needles m. p. 113~114°C.

Found: C, 37.41; H, 1.19. Calcd. for $C_{13}H_7OBr_3$: C, 37.28; H, 1.99%.

The benzene eluent gave 0.6 g. of IX, 0.7 g. of IV and 0.1 g. of XI, successively. From the chloroform eluent, 0.1 g. of I was obtained, acompanied with a resinous substance. Alcohol eluent did not give any crystals.

In this case, the alkaline washing solution gave no crystals by acidification.

ii) To a solution of 4.7 g. of II in 7.5 cc. of acetic acid, 20 g. of bromine dissolved in 3.5 cc. of acetic acid was added. About 3/4 vols. of bromine solution was added at 40~75°C for a period of 1.5 hrs., then yellow crystals were separated out. These were dissolved again in a dark red solution when the rest of the bromine was added at 85°C. The mixture was further heated for 6 hr. until there was no liberation of hydrogen bromide. After cooling, the yellow crystals (a) thereby separated were collected by filtration. There were digested with benzene and an insoluble portion (b) was filtered off. The filtrate was shaken with 10% sodium carbonate solution and the aqueous layer was combined with (b). This alkaline solution, after being treated with active carbon, was acidified to give 0.27 g. of X, m. p. 200~205°C. The benzene layer, after removal of the solvent, left a crystalline residue which was recrystalized from alcohol or ether to give 4.5 g. of IX.

The mother liquor of (a), treated as in a, i), gave the following crystals: 1.0 g. of IX, 0.44 g. of X, 0.1 g. of XI, 0.2 g. of XII and 0.15 g. of XIII.

iii) To a solution of 4.7 g. of II in 7.5 cc. of acetic acid and 2.5 cc. of acetic anhydride, 20 g. of bromine dissolved in 3.5 cc. of acetic acid was added as in the above case. In this case, no crystals were separated during the reaction. On cooling, 5.3 g. of III was obtained and this was decomposed to 3.1 g. of IV. After being treated as in a, i), 0.2 g. of X was obtained from the 10% sodium carbonate soluble part of the mother liquor. The chromatographic separation of the neutral substance gave successively the following crystals: 0.35 g. of IX, 0.40 g. of IV, 0.15 g. of XX and 0.1 g. of I.

The recrystallization of XX from methanol gave yellow needless of m. p. 144~145°C.

Found: C, 60.29; H, 3.58. Calcd. for $C_{13}H_9OBr$: C, 59.79; H, 3.47%.

Action of Hydrazine on 4-Bromo-2-phenyltropone (IV).—A mixture of 0.2 g. of IV, 0.2 cc. of 80% hydrazine hydrate and 2 cc. of alcohol was refluxed on a water bath for 30 min., and the crystals that separated out were collected by filtration. Recrystallization from alcohol gave orange yellow needles of VI, m. p. 210°C.

A mixture of 0.08 g. of VI, 0.3 cc. of 6 N sodium hydroxide solution and 1.6 cc. of alcohol was refluxed for 3 hr. The alcohol was distilled off, the residue was acidified, and the precipitate thereby formed was recrystallized from methanol to give yellow needles of m. p. 135~136°C. This was found to be identical by mixed fusion with the authentic sample of 5-bromo-3-phenyltropolone (V)[10].

Bromination of 4-Bromo-2-phenyltropone (IV).—To a solution of 2.6 g. of IV in 40 cc. of acetic acid, 4.8 g. of bromine in 5 cc. of acetic acid was added and this mixture was allowed to stand over night. The crystals that separated out were collected and recrystallized from alcohol and benzene to give colorless prisms of XIV, m. p. 121~122°C (decomp.). Yield, 4.7 g.

Found: C, 27.26; H, 1.68. Calcd. for $C_{13}H_9OBr_5$: C, 26.89; H, 1.56%.

The filtrate was evaporated under reduced pressure and the residue, after being treated with water, was extracted with benzene. After removal of solvent, the extract left a crystalline residue which was recrystallized from alcohol to give yellow crystals of m. p. 110~113°C. This showed no depression of melting point on admixture with IX obtained in the above cases.

Decomposition of 4,4,5,6,7-Pentabromo-2-phenyl-2-cyclohepten-1-one (XIV).—i) By heating 0.15 g. of XIV in an oil bath at 120~130°C for 10 min., XIV melted with effervescence and underwent decomposition to an oil. After cooling, this was dissolved in benzene and washed with a diluted sodium hyposulfate solution. The benzene layer, after the removal of solvent, left a crystalline residue that was recrystallized from alcohol to give yellow crystals of m. p. 116~117°C, undepressed on admixture with IX. Yield, 0.05 g.

ii) A mixture of 0.58 g. of XIV, 0.6 cc. of pyridine and 0.3 g. of benzene was heated for 15 min. at 55~60°C. This was diluted with water and extracted with benzene. The extract after evaporation, left yellow crystals. Recrystallization from alcohol gave yellow needles of m. p. 112~113°C, undepressed on admixture with XII obtained above.

Action of Alkali on 4, 7-Dibromo-2-phenyltropone (IX).—A mixture of 0.06 g. of IX, 1.5 cc. of alcohol and 1 cc. of 2 N sodium hydroxide solution was refluxed on a water bath for 30 min. After the removal of solvent it was acidified and extracted with benzene. The benzene layer gave colorless crystals that were recrystallized from cyclohexane to give XV as colorless needles of m. p. 173~175°C.

Found: C, 56.75; H, 3.47. Calcd. for $C_{13}H_9OBr$: C, 56.34; H, 3.27%.

Action of Alkali on 4, 5, 7-Tribromo-2-phenyltropone (XII).—A mixture of 0.05 g. of XII, 0.7 cc. of alcohol and 0.5 cc. of 2 N sodium hydroxide was refluxed on a water bath for 15 min. This, treated as in the above case, gave colorless needles of XVIII, m. p. 180~181°C.

Found: C, 43.76; H, 2.15. Calcd. for $C_{13}H_8O_2Br_2$: C, 43.76; H, 2.26%.

Action of Sulfuric Acid on 4, 5, 7-Tribromo-2-phenyltropone(XII).—A solution of 0.25 g. of XII and 2.5 cc. of 75% sulfuric acid was heated for 10 hr. at 170~180°C. This was poured into about 20 cc. of ice water, and thereby precipitated crystals were collected. The crystals, after being dried, were washed with benzene and dissolved in 10% sodium carbonate solution. After being treated with active carbon, the solution was acidified and 0.1 g. of yellow crystals of d. p. ca 190°C were obtained. This was treated with a 2.8% diazomethane-ether solution and the thereby obtained oily product was dissolved in a mixture of benzene and petroleum ether. This was chromatographed on a silica-gel column and gave yellow crystals of m. p. 128~129°C. Recrystallization from methanol gave yellow needles of m. p. 131~132°C, which showed no depression with authentic sample of XVI.

Thermal Decomposition of 4, 7, *x*, *x*-Tetrabromo-2-phenylsuberenone(XIII).—By heating 0.12 g. of XIII on an oil bath for 20 min. at 150~160°C, it melted and underwent decomposition. The fused mass, dissolved in benzene, was washed sufficiently with water. The benzene solution was passed through a silica-gel column and yellow needles of m. p. 117~118°C were obtained. These showed no depression on admixture with IX obtained above.

Bromination of *x*-Bromo-2-phenyltropone(XX).—To a solution of 0.06 g. of XX in 1 cc. of acetic acid, 0.06 g. of bromine was added and the whole was heated on a water bath for 30 min. Then this was diluted with water and extracted with chloroform. After the removal of solvent, this gave colorless crystals of m. p. 135~140°C, which were recrystallized from benzene or alcohol to give pale yellow needles, m. p. 141~142°C.

Found: C, 45.94; H, 2.36. Calcd. for $C_{13}H_9OBr_2$: C, 45.93; H, 2.36%.

The author takes this opportunity to express his sincere gratitude to professor Tetsuo Nozoe of Tohoku University for his guidance and encouragement throughout the course of this work, to Professor Shuichi Seto, Professor Yoshio Kitahara and Assistant Professor Toshio Mukai for much helpful advice, and to assistant Professor Katsuo Kikuchi of Yamagata University for his kind discussion and helpful advice.

Department of Chemistry
Faculty of Liberal Arts and Science
Yamagata University
Koshirakawa-machi, Yamagata

---

*Stereoregular Polymers. VI. Copolymerization with the Ziegler Catalysts**

By Shunsuke MURAHASHI, Shun'ichi NOZAKURA and Masao SUMI

(Received January 5, 1960)

Binary copolymerizations of α-olefins, which are known individually to form isotactic polymers with the Ziegler catalyst, were investigated to see if stereospecific copolymerization could be seen to yield crystalline copolymers.

Although there are a number of reports by Natta and his collaborators on copolymerization of propylene and ethylene with complex catalysts of the Ziegler type[1], one of the comonomers was always ethylene and this has substantially no concern with the assymmetricity on the polymer. Besides, ethylene has markedly higher activity towards the anionic polymerization as compared to propylene, and their efforts were apparently directed towards the study of amorphous copolymers.

In the present work, the situation was distinctly different from that in the case of Natta. Particular attention was directed to the formation of crystalline copolymers, because it was believed that crystalline copolymers could be produced if the isotactic orientation had been held during copolymerization**.

The monomers chosen here are styrene, allyl- and 3-butenyl-trimethylsilane, which are known by the present authors to yield the corresponding isotactic polymers by themselves and to have the similar reactivity towards the Ziegler catalyst[3,4].

The following abreviations will be used in

* Presented at the 8th. Annual Meeting of Society of Polymer Science (Japan), May 30, 1959.

1) G. Natta et al., *Chim. e ind.*, **39**, 733, 743, 825 (1957); **40**, 717, 896 (1958).

** In the course of our study, Reding has reported in a brief note the cocrystalline copolymers of 3-methyl-1-butene and 4-methyl-1-pentene, which were found to be surprisingly highly crystalline over the entire range of compositions.[2]

2) F. P. Reding and E. R. Walter, *J. Polymer Sci.*, **37**, 555 (1959).

3) S. Murahashi, S. Nozakura and M. Sumi, This Bulletin, **32**, 670 (1959).

4) G. Natta, G. Mazzanti, P. Longi and F. Bernardini, *J. Polymer Sci.*, **31**, 181 (1958).

TABLE I. COPOLYMERIZATION OF STYRENE AND ALKENYLTRIMETHYLSILANES WITH $AlEt_3$-$TiCl_4$ CATALYST AT 70°C

Combination of monomers	Monomer, mole		Catalyst, millimole			*n*-Heptane ml.	Time hr.	Polymer yield, %	Composition, mole % of silane*	
	Silane	St	$AlEt_3$	$TiCl_4$	Al/Ti ratio				Monomer	Polymer
St–VSi	0.0298	0.0290	4.30	1.79	2.40	20	7	16.0	50.48	1.94
St–ASi	0.0176	0.0724	5.61	2.26	2.48	20	3.5	13.2	19.56	6.19
	0.0166	0.0173	3.77	1.42	2.65	15	6	23.3	48.97	18.87
	0.0615	0.0165	5.87	2.21	2.66	20	2	10.1	78.85	66.62
St–BSi	0.0146	0.0624	5.70	2.21	2.58	20	3	15.3	18.96	10.64
	0.0378	0.0379	5.87	2.32	2.53	20	3	20.5	49.93	42.71
	0.0650	0.0170	6.87	2.63	2.60	20	4	27.4	79.27	70.21

* Copolymer composition was determined by the gravimetric analysis of silicon.

TABLE II. COPOLYMERIZATION OF ASi AND BSi WITH $AlEt_3$-$TiCl_4$ CATALYST AT 70°C

Monomer, mole		Catalyst, millimole			*n*-Heptane ml.	Time hr.	Polymer yield, %	Composition, mole % of silane	
ASi	BSi	$AlEt_3$	$TiCl_4$	Al/Ti ratio				Monomer	Polymer
0.0306	0.0289	5.35	2.21	2.42	20	3.5	22.1	51.43	8.74
0.0735	0.0192	6.58*	2.63	2.50	25	3.5	9.9	79.29	48.72

* A solution of $AlEt_3$(3.86 millimole) was added to a boiling *n*-heptane solution of $TiCl_4$ (2.63 millimole) and after cooling it the precipitated mass was washed with *n*-heptane, and then $AlEt_3$(2.72 millimole) was added.

this paper for convenience:

Vinyltrimethylsilane	VSi
Allyltrimethylsilane	ASi
3-Butenyltrimethylsilane	BSi
Styrene	St

Results and Discussion

Polymerization. — Binary copolymerizations were carried out using these monomers with $AlEt_3$–$TiCl_4$ catalyst at 70°C. The results were summarized in Tables I and II.

VSi, the first member of alkenyltrimethylsilane, was found to enter into copolymerization with great difficulty as anticipated from the case of its homopolymerization with the same catalyst, and hence a further study was omitted from our program.

As is shown in Tables I and II, the compositions of the crude polymers varied with monomer compositions and copolymerization appeared at first to have been realized. The apparent monomer reactivities towards the Ziegler catalyst was found to be in the order of St > BSi > ASi ≫ VSi.

Fractionation of the Crude Polymers. — As these copolymers showed some degree of crystallinity on X-ray diffraction patterns, fractionations were carried out with a series of solvents selected from the knowledge of the solubility characteristics of the individual crystalline homopolymers. (Table III).

TABLE III. SOLUBILITY CHARACTERISTICS OF ISOTACTIC POLYMERS OF St, ASi AND BSi

	n-Heptane	Toluene	Decalin
Poly(BSi)	Sol. in hot	Sol.	Sol.
Poly(St)	Insol.	Sol. in hot	Sol.
Poly(ASi)	Insol.	Insol.	Sol. in hot

Thus, a St–BSi copolymer was fractionally extracted with a series of solvents: acetone, ether, methylethylketone, *n*-heptane and toluene under reflux in this order. The results are shown in Table IV in which was included, for reference, the results of model fractional extractions of an artificial mixture of two homopolymers, showing practically a complete separation of the crystalline parts into the original components. The acetone-, ether- and methylethylketone-soluble fractions of the copolymer are all found to give almost amorphous X-ray diffraction patterns, while both the *n*-heptane- and the toluene-fraction are found to give crystalline patterns. Analytical data of the last two fractions show that the *n*-heptane-fraction is closely similar to isotactic poly(BSi) with a minor amount of St, while the toluene-fraction is similar to isotactic poly(St) with a minor amount of BSi. The amount of the minor ingredient (St) in the former fraction showed no appreciable change even after repeated purification, that is, extraction with methylethylketone followed by dissolution of the residue into hot *n*-heptane and reprecipitation by cooling the solution. Therefore, it

TABLE IV. FRACTIONATION OF St-BSi COPOLYMERS

Polymers	Composition of the crude polymer	Acetone-extract: Sol. in cold		Acetone-extract: Sol. in hot		Ether extract		M. E. K extract		*n*-Heptane extract		Toluene extract	
	mole % of silane	%	mole % of silane	%	mole % of silane	%	mole % of silane	%	mole % of silane	%	mole % of silane	%	mole % of silane
Copolymer	10.64	23.1	7.35	5.5	35.03	3.3	49.59	1.8	—	3.3	76.72	53.7	2.76
Copolymer	42.25	19.9	18.68	10.1	49.23	18.6	67.62	3.1	14.48	19.6	89.57	22.4	3.81
Copolymer	70.21	5.4	40.73	8.9	67.22	13.2	76.46	0.8	—	53.0	90.61	4.6	13.53
Mixture	40.64	5.8	22.75	1.2	—	6.2	94.60	0.7	—	29.4	92.79	43.7	0.19
X-ray evidence		amorph.		amorph.		amorph.		amorph.		cryst.		cryst.	

TABLE V. FRACTIONATION OF St-ASi COPOLYMERS

Polymers	Composition of the crude polymer	Acetone extract		Ether extract		*n*-Heptane extract		M. E. K. extract		Toluene extract		Toluene residue		Decalin extract	
	mole % of ASi	%	mole % of ASi	%	mole % of ASi	%	mole % of ASi	%	mole % of ASi	%	mole % of ASi	%	mole % of ASi	%	mole % of ASi
Copolymer	6.19	9.0	19.84	1.3	—	1.4	—	2.5	—	78.1	1.52	1.2	—		
Copolymer	18.89	13.3	33.76	4.8	51.57	1.2	53.91	0.9	18.02	63.4	2.48	10.6	77.44		
Copolymer	66.62	16.1	—	11.5	—	4.2	—	0.9	—	17.9	9.19			0.79	85.07
Micture	19.39	6.4	52.14	2.4	—	1.0	—	0.5	—	68.8	0.22	11.8	95.50		
X-ray evidence		amorph.		amorph.		amorph.		amorph.		cryst.		cryst.			

TABLE VI. FRACTIONATION OF ASi-BSi COPOLYMERS

Polymers	Crude polymers Si, %*	Acetone extract %	Ether extract %	*n*-Heptane extract %	*n*-Heptane extract Si, %*	Decalin extract %	Decalin extract Si, %*
Copolymer	22.13	20.3	45.7	18.6	—	6.6	—
Copolymer	23.13	15.9	29.3	14.7	21.0	31.0	24.0
X-ray evidence		amorph.	amorph.	cryst.		cryst.	

* The theoretical Si content of BSi and ASi are 21.90 and 24.59%, respectively. The analysis of the fractions was carried out on a microscale.

was concluded that the *n*-heptane-fraction might not be a mixture but a genuine copolymer. It was also found that the same was held to be the case by the latter fraction.

An additional evidence for the copolymeric character of the two crystalline fractions is the change of their compositions with the fed monomer compositions. By increasing the silane content in monomer composition, the silane content of both the *n*-heptane- and the toluene-fraction increased.

Thus, it was confirmed that the crystalline portion of the St–BSi polymer consisted of two kinds of crystalline copolymers of quite different compositions, that is, both a crystalline copolymer with a predominant portion of one component and another crystalline copolymer of another predominant component were produced side by side.

Similar results hold in the copolymerization with the combinations of St–ASi and ASi–BSi (Tables V and VI), though in the latter case the pair selected was of very similar monomers.

From these results, one might postulate that the catalyst surface might possess two kinds of isotactically active sites for each monomer. One kind of the sites is more favorable to polymerize one monomer than other monomers and this is true of other site. More generally, one might say that the catalytic surface possesses several different kinds of active sites suitable for some kinds of monomers. Copolymerization would have taken place on such sites, leading to the crystalline copolymers of different compositions.

An alternative interpretation might also be possible where the nature of an active site may not have been set specifically beforehand at the time when the catalyst is formed, and may become specific after the first attack of a monomer molecule on the catalyst site.

It may not be easy to distinguish which of the two alternative mechanisms is more favorable in our case, but the following experimental results seems to show that the former mechanism might be more favorable. It was demonstrated in one experiment where a small amount of one monomer (about 1millimole) was added preliminarily to the catalyst at 70°C, and after 20 minutes of polymerization a mixture of two monomers St and ASi (20 millimole : 20 millimole) was then added. Two runs that were carried out by the preliminary addition of St in one case and ASi in another case resulted in practically the same distribution of fractions in the extraction experiment.

X-Ray Examination of the Crystalline Fractions.—In the case of the copolymerization of St and BSi, for instance, the *n*-heptane- and the toluene-fraction showed some degree of

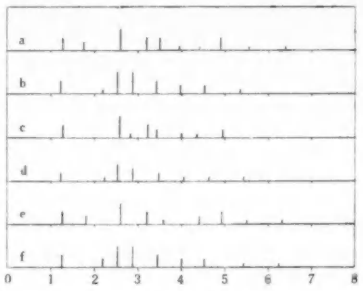


Fig. 1. X-Ray spectra of ASi-St copolymers. (R=4.45 cm.)
a) Copolymer : silane, 77.4 mole %.
b) Copolymer : silane, 9.1 mole %.
c) Artificial mixture : silane, 77 mole %.
d) Artificial mixture : silane, 9.2 mole %.
e) Isotactic poly(allylsilane).
f) Isotactic polystyrene.

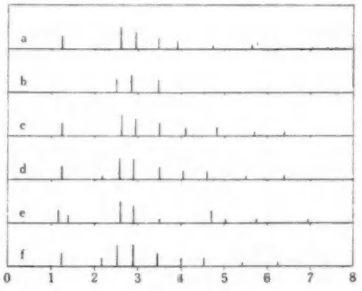


Fig. 2. X-Ray spectra of BSi-St copolymers. (R=4.45 cm.)
a) Copolymer : silane, 89.6 mole %.
b) Copolymer : silane, 13.5 mole %.
c) Artificial mixture : silane, 73 mole %.
d) Artificial mixture : silane, 23 mole %.
e) Isotactic poly(3-butenylsilane).
f) Isotactic polystyrene.

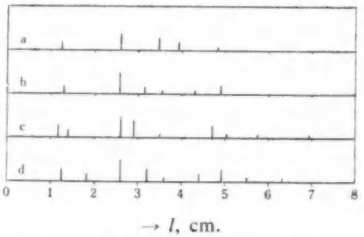


Fig. 3. X-Ray spectra of ASi-BSi copolymers. (R=4.45 cm.)
a) Copolymer : the *n*-heptane-fraction.
b) Copolymer : the decalin-fraction.
c) Isotactic poly(3-butenylsilane).
d) Isotactic poly(allylsilane).

crystalline patterns on X-ray diagrams. X-Ray powder patterns of these two fractions were compared with those of two homopolymers and of the artificial mixtures of similar compositions. The same comparison was also made in other cases. X-Ray spectra of St–ASi copolymers, St–BSi copolymers, and ASi–BSi copolymers were shown in Figs. 1, 2 and 3, respectively***.

From the close inspection of these results, it was found that in some cases the patterns closely resemble those of artificial mixtures of the similar compositions (Figs. 1a and 1b), some had a new kind of additional reflections over those of artificial mixtures or a superposition of homopolymers (Figs. 2a and 3a), and others showed only very diffuse reflections of one homopolymer (Figs. 2b and 3b). These facts might suggest the possible existence of constitutional varieties in copolymers with the change of monomer species, that is, one copolymer might have a structure of block copolymer and another might have a structure of random copolymer. For instance, the decalin-fraction (Fig. 1a) of St–ASi copolymer has a rather high crystallinity which is comparable to that of poly(ASi) and shows a diffraction pattern that is closely similar to an artificial mixture of the same composition, suggesting a block structure. If it were a random copolymer, incorporation of about 20% of St units, as is the case, would almost destroy the crystallinity of the copolymer.

Experimental

Materials.—A commercial styrene was purified according to the usual method[5].

Vinyl-, allyl, and 3-butenyl-trimethylsilane are prepared and purified just as described in the previous paper[4].

$AlEt_3$ and $TiCl_4$ are the products of Mitsui Chem. Ind. Co. and Osaka Titanium Co., respectively and were used as about 30 and 21% solutions in *n*-heptane.

Polymerization.—Polymerization was carried out without agitation in an apparatus described in the previous paper[6]. The solution of $AlEt_3$ was added first to *n*-heptane in the apparatus under nitrogen, and mixed, and then the solution of $TiCl_4$ was dropped slowly along the wall of the vessel. After ten minutes, the mixed monomers were added and polymerization was carried out without agitation.

Polymerization was stopped by the addition of methanol (2 ml.) and concentrated hydrochloric acid (0.3 ml.), and the mixture was poured into a large amount of methanol. After being kept overnight, the polymer was collected, washed with methanol and dried.

Fractionation of the Crude Copolymers.—Hot extraction was carried out at the boiling temperature of the solvents, using a modified Soxhlet extraction apparatus. Complete extraction with one solvent generally took about seven to ten hours. After the completion of one extraction, the extraction cup with the residual polymer in it was dried, and was subjected to the next extraction. Each extract was concentrated to about 50 ml. and poured into a large amount of methanol, and the precipitate was collected, washed and dried.

Purification of the crystalline fraction, for example, the *n*-heptane fraction of St–BSi copolymer was carried out as follows: the fraction (Si, 20.08%) was extracted with methyl ethyl ketone, and dissolved in hot *n*-heptane. The hot solution, after being filtered through a glass filter, was poured into a large amount of methanol. The precipitated polymer was collected, washed and dried in vacuum at about 100°C (Si, 19.92%). The toluene fraction (Si, 1.80%) of the polymer was also purified by extraction with *n*-heptane followed by dissolution in toluene and reprecipitation with methanol, giving the silicon content of 1.40 and 1.48%, successively.

Preparation of Artificial Mixture of Homopolymers.—Generally, a mechanical mixture of two kinds of polymers was dissolved in a hot solvent common to them and poured quickly into a large amount of methanol and then the precipitate was collected.

Determination of Copolymer Composition.—The composition of copolymers was determined by the gravimetric analysis of silicon. Analysis was carried out by wet digestion with fuming sulfuric acid in a platinum crucible. The ignited silica was ascertained by treating it with 60% hydrofluoric acid.

X-Ray Study.—All samples were annealed at 150~160°C for 2 hr. The diffraction patterns were taken using nickel filtered Cu$K\alpha$ radiation. The film distance was 4.45 cm.

Summary

Binary copolymerization of styrene, allyltrimethylsilane and 3-butenyltrimethylsilane was investigated using the Ziegler catalyst at 70°C. It was found to be possible to separate the crystalline part of copolymers by fractional extraction into two kinds of genuine crystalline copolymers, each of which has an extremely different composition from the other.

The nature of the active sites on the catalytic surface was discussed in order to elucidate the above fact.

The constitution of the individual crystalline fractions appeared to vary case by case, i.e., the structure of blockcopolymer might be assigned to some cases and random copolymers to some other cases.

Department of Polymer Science
Faculty of Science
Osaka University
Nakanoshima, Kita-ku, Osaka

*** Since the exact copolymer composition was not obvious in the case of ASi-BSi copolymer because of the small difference of silicon content between ASi and BSi units, comparison with artificial mixtures was not indicated in the list.

5) E. R. Blout and H. Mark, "Monomers", Interscience Publisher (1944).

6) S. Murahashi, S. Nozakura, M. Sumi and K. Hatada, This Bulletin, 32, 1094 (1959).

The Aconite Alkaloids. XXXIV[1]*. On Lucidusculine. Part IV*[2]

By Takashi AMIYA

(Received January 6, 1960)

In previous reports[3,4] the partial formula of lucidusculine (I), $C_{24}H_{35}O_4N$, has been summarized as[5],

$$C_{16}H_{20}\begin{cases} CH_2OH \\ OCOCH_3 \\ CH(OH) \text{ (six-membered ring)} \\ NC_2H_5 \\ C{=}CH_2 \end{cases}$$

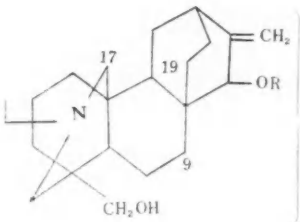
(I)

Moreover, luciculine (II), $C_{22}H_{33}O_3N\cdot H_2O$, which is yielded from compound I by hydrolysis and hence is a sort of trihydric amino alcohol, has been shown to possess an allyl alcohol moiety in its six-membered ring[6].

On the other hand, on dehydrogenation of compound I, phenanthrene and its homologues were obtained by Suginome and Kakimoto[7] in this laboratory.

Under the same conditions as in the case of isomerization of compound II to isoluciculine (III)[6], compound I was found not to isomerize. Instead, a certain compound IV was formed. This compound did not crystallize. It seemed to be a dihydro compound. In the infrared spectrum, compound IV gave no band characteristic of a ketone carbonyl group in a six-membered ring near 5.8 μ but it did give a band characteristic of an ester carbonyl group. No absorption band near 11 μ, indicative of a terminal methylene group, was observed in the spectrum. On hydrolysis of compound IV by alcoholic potassium hydroxide, an amorphous base V was obtained. In the infrared region, this base showed no band near 5.7~5.8 μ, indicating that carbonyl group, either ester or ketone, was quite absent. The spectrum also showed that no band characteristic of a terminal methylene group near 11 μ existed. This fact may show that the acetylated hydroxyl group in compound I is the secondary one involved in the allyl alcohol moiety mentioned above. The above hydrogenation seems to proceed similarly to that of Shimoburo base I diacetate[8].

On the basis of these experimental results as well as in view of the structures of other aconite alkaloids, the partial formula of compounds I and II is now represented in Fig. 1.



—OH (a secondary one in a six-membered ring)

(I) R=Ac
(II) R=H

Fig. 1.

Lucidusculine has been assumed to have a hexacyclic structure[9]. Considering the structures of hypogenavine[10], kobusine[11], ignavine[12], hetisine[13] and songorine[14], the existence of the C-17—C-19 or C-17—C-9 bond seems to be reasonable. Moreover, to complete the structure of compound I, it is necessary to confirm the position of an indefinite secondary hydroxyl group (see Fig. 1).

Experimental

Hydrogenation of Lucidusculine (I) with Pd-C. —Lucidusculine (0.3 g.) was dissolved in 15 cc. of methanol and hydrogenated with 10% palladium-carbon (0.2 g.) at room temperature for 4.5 hr. After removal of palladium-carbon, the solution was evaporated to dryness. On treatment with methanol, the residue was not crystallized. The infrared absorption spectrum in nujol had no peak attributable to a terminal methylene group but

1) This constitues a part of a series entitled "The Aconite Alkaloids" by H. Suginome. Part XXXIII: cf. Ref. 4).
2) XXXIII of this series corresponds to part III of "On Lucidusculine".
3) T. Amiya, This Bulletin, **30**, 677 (1957).
4) T. Amiya, ibid., **33**, 644 (1960)..
5) The partial formula of I was presented in a previous report[3], as $C_{17}H_{20}$ erroneously, but this should be corrected.
6) T. Amiya, This Bulletin, **32**, 421 (1959).
7) H. Suginome and S. Kakimoto, ibid., **32**, 352 (1959).
8) T. Sugasawa, *Pharm. Bull.*, **4**, 6 (1956).
9) H. Suginome, T. Amiya and T. Shima, This Bulletin, **32**, 824 (1959).
10) S. Sakai, *Chem. Pharm. Bull.*, **6**, 448 (1958).
11) M. Natsume, ibid., **7**, 539 (1959).
12) E. Ochiai and T. Okamoto, ibid., **7**, 556 (1959).
13) A. J. Solo and S. W. Pelletier, *J. Am. Chem. Soc.*, **81**, 4439 (1959).
14) E. Ochiai, T. Okamoto, S. Sakai, T. Sugasawa and T. Onouchi, *Chem. Pharm. Bull.*, **7**, 542 (1959).

showed the presence of hydroxyl groups (2.92~2.94 μ) and an acetoxyl group (5.80 and 8.10 μ).

Hydrolysis of Compound IV. — A solution of compound IV (0.05 g.) and potassium hydroxide (0.15 g.), dissolved in 2 cc. of ethanol and a drop of water, was heated for 1.5 hr. After removal of the solvents, the residue was washed with water. On treatment with methanol, it was not crystallized. The infrared absorption spectrum in nujol had no peak attributable to a terminal methylene group but showed the presence of hydroxyl groups (2.97 μ).

The author wishes to express his sincere thanks to Professor Harusada Suginome, President of Hokkaido University, for his unfailing kindness in encouraging this work.

Department of Chemistry
Faculty of Science
Hokkaido University
Sapporo

Halogen Substitution in Liquid Sulfur Dioxide. I. Formation of Alkyl Halides from Alcohols and Halogens

By Niichiro TOKURA, Ritsuro TADA and Ryoichi IGARASHI

(Received December 28, 1959)

The writers clarified the point previously[1] that ketoximes undergo rearrangement by the action of a halogen in liquid sulfur dioxide at room temperature. Scheme 1 was tentatively assigned to the mechanism of the rearrangement. The solvation by the solvent (liq. SO_2) may favor the ionization of the bromine molecule.

Sheme 1

$$Br_2 \overset{\text{liq. } SO_2}{\rightleftharpoons} Br^+ + Br^- \quad (1)$$

$$\begin{matrix} R \\ R' \end{matrix}\!\!>\!C{=}N{-}OH \xrightarrow{Br^+} \begin{matrix} R \\ R' \end{matrix}\!\!>\!C{=}N^+ + BrOH \quad (2)$$

$$\begin{matrix} R \\ R' \end{matrix}\!\!>\!C{=}N^+ + Br^- \longrightarrow \begin{matrix} R \\ Br \end{matrix}\!\!>\!C{=}N{-}R' \xrightarrow{H_2O} R{-}\underset{\underset{O}{\|}}{C}{-}NHR' \quad (3)$$

$$BrOH + SO_2 \longrightarrow HBr + SO_3 \quad (3)$$

From the structural similarity between alcohols R–C–OH and oximes R–C=H–OH, it is expected that all alcohols have a similar behavior towards bromine in liquid sulfur dioxide and, consequently, the reaction will proceed to form the corresponding alkyl halides. (Scheme 2).

Scheme 2

$$ROH + Br^+ \longrightarrow R^+ + BrOH \quad (5)$$

$$R^+ + Br^- \longrightarrow RBr \quad (6)$$

$$BrOH + SO_2 \longrightarrow HBr + SO_3 \quad (7)$$

1) N. Tokura, R. Asami and R. Tada, *J. Am. Chem. Soc.*, **76**, 3135 (1957).

The borderline nature[2] of liquid sulfur dioxide as a solvent was first recognized by Ingold[2] and, if the reaction follows the SNI mechanism, Eq. 5 being a slow step, the ease of the reaction of alcohols will be in the order of just tertiary, then, secondary, and finally, primary.[3]

A direct substitution of an alcoholic hydroxyl by a bromine atom has been reported to be the case, though only in tertiary alcohols, but this method[4] is not used because of its low yield. Substitution of the hydroxyl in liquid sulfur dioxide by a halogen has already been carried out[5], but with such halogenating reagents as thionyl halides, phosphorus trihalides and pentahalides, liquid sulfur dioxide being used as a solvent because of its excellence.

In expectation of obtaining alkyl halides by direct application of a halogen in liquid sulfur dioxide at room temperature, the lower members of aliphatic and alicyclic alcohols were submitted to the reaction with bromine with the results summarized in Table I.

The total amount of the products increased in the order of primary, secondary and tertiary alcohols. Only small amounts of alkyl halides were obtained from primary alcohols. Higher boiling halides were produced by further

2) C. K. Ingold, *Proc. Chem. Soc.*, 279 (1957).
3) C. K. Ingold, "Structure and Mechanism in Org. Chem.", Cornell Univ. Press, Ithaca, N.Y. (1955), p. 317.
4) C. Hell and F. Urech, *Ber.*, **15**, 1249 (1882).
5) For example, R. B. Woodward et al., *J. Am. Chem. Soc.*, **76**, 5256 (1954), used thionyl chloride in liq. SO_2 for the substitution of hydroxyl with chlorine in their synthetic course to lysergic acid, one of the ergot alkaloids.

TABLE I. REACTION OF BROMINE AND ALCHLOLS IN LIQ. SO_2
Alcohol: 0.4 mol., Bromine: 0.4 mol. (64 g.)
Liq. SO_2: 100 mol., Time: 5 hr. at room temperature

Alcohol ROH	R	Ethyl	*n*-Propyl	Iso-propyl	*n*-Butyl	Iso-butyl	*sec*-Butyl	*tert*-Butyl	Iso-amyl
	g.	18.4	24.0	24.0	30.0	30.0	30.0	30.0	35.0
Product: Total	g.	3.0	3.0	30.0	8.0	17.0	60.0	69.0	4.0
Product: Total	Yield weight %	3.6	3.4	34.1	8.5	18.1	63.9	73.3	4.0
Product: Alkyl bromide RBr	b.p., °C	37~38	68~70	59~60	99~101	88~91	90~92	71~73	120~121
Product: Alkyl bromide RBr	g.	2.0	2.5	19.0	5.0	1.5	18.0	13.0	3.0
Product: Alkyl bromide RBr	Yield, %	5.3	5.1	38.6	7.2	2.7	32.5	23.4	6.7
Product: Alkyl bromide RBr	n_D^{25}	1.425	1.435	1.423	1.440	1.438	1.438	1.430	1.443
Product: Alkyl bromide RBr	Br % found.	71.5	63.0	62.5	57.9	58.1	56.1	57.3	53.0
Product: Alkyl bromide RBr	Br % calcd.	73.0	64.9	64.9	58.3	58.3	58.3	58.3	52.9
Product: Alkyl bromide RBr	$RCONHC_6H_5$ m.p., °C	104	91~92	101~10	62~63	108~109	107~108	127~128	107~108
Higher boiling fractions				b.p. 42~43 (20 mmHg) 7.0 g.		b.p. 41~44 (10 mmHg) 9.0 g.	b.p. 41~44 (12 mmHg) 30.0 g.	b.p. 46~51 (20 mmHg) 31.0 g.	
						b.p. 90~99 (10 mmHg) 5.0 g.	b.p.90~100 (10 mmHg) 2.0 g.	b.p.80~110 (20 mmHg) 4.0 g.	

elimination, dimerization and bromination. In all experimants sulfuric acid separated out on the bottom of the reaction vessel. The highest yield was 38.6% of isopropyl bromide, but it was accompanied by a hexyl bromide with a carbon number twice as large as the original alkyl and most of the butyl bromides were contaminated with butylene dibromides. (Scheme 3)

Scheme 3

$$R_1R_2\cdot(R_3CH_2)\cdot C{-}OH \xrightarrow{Br^-} A^+ + BrOH \quad (8)$$

$$A^+ ; R_1R_2\cdot(R_3CH_2)\cdot C^+$$

$$A^+ + Br^- \longrightarrow R_1R_2(R_3CH_2)\cdot C\cdot Br \quad (9)$$

$$A^+ \longrightarrow R_1R_2(R_3CH_2)\cdot C{=}CH\cdot R_3\ (B) + H^+ \quad (10)$$

$$BrOH + SO_2 \longrightarrow HBr + SO_3 \quad (11)$$

$$R_1R_2(R_3CH_2)\cdot C{-}OH + SO_3 \longrightarrow A^+ + HSO_4^- \quad (12)$$

$$H^+ + HSO_4^- \longrightarrow H_2SO_4 \quad (13)$$

$$B + A^+ \longrightarrow R_1R_2\cdot C{-}R_3CH{-}C\cdot R_1R_2\cdot(CH_2R_3)$$

$$\xrightarrow{Br^-} R_1R_2{-}C{-}R_3CH{-}CR_1R_2(CH_2R_3)^* \quad (14)$$

$$B + Br_2 \longrightarrow R_1R_2C{-}\underset{\displaystyle Br}{\underset{|}{C}}H\cdot R_3 \quad (15)$$

where R_1, R_2, R_3=H or CH_3
* more isomers are possible.

Thus, as might be expected, elimination reaction prevails to a considerable extent in liquid sulfur dioxide. The rather complicated feature of the reaction will be understood if the sulfur trioxide produced by the secondary step is assumed to play an important role in the elimination reaction and the dimerization.

$$RBr \rightleftharpoons R^+ + Br^- \longrightarrow \text{Olefin} + H^+ Br^-$$

$$\text{Olefin} \xrightarrow{+Br_2} \text{dibromide}$$

The equilibrium tends to the formation of the dibromide. It is further supported by the fact that phenethyl chloride and bromine in liquid sulfur dioxide afford styrene, polystyrene and styrene dibromide[6].

As the reaction proceeds, sulfur trioxide, formed according to Eq. 7, will favor the formation of the alkyl cation followed by elimination to olefin and sulfuric acid thus produced separates out owing to its slight solubility in liquid sulfur dioxide. Since this is an irreversible reaction, the reaction reaches a final state when the whole acid is precipitated.

In the case of the Beckmann rearrangement, the formation of a nitrogen cation or its ion-pair is the main reaction and no elimination occurs, the secondarily formed sulfur trioxide bearing a lesser meaning.

In the bromination of alcohols by bromine in liquid sulfur dioxide, the decisive factor is the SNI tendency of both alcohols and the solvent. The larger amounts of reaction products from secondary and tertiary alcohols are well consistent with the above consideration.

The writers have no evidence to rule out the participation of hydrogen bromide afforded by Eq. 7 in the bromine substitution reaction. An exact elucidation of the reaction needs further examination.

6) E. D. Hughes, C. K. Ingold and A. D. Scott, *J. Chem. Soc.*, **1937**, 1271.

Table II. Effect of the amount of bromine on the yield and action of hydrogen bromide on isopropyl alcohol

Isopropyl alcohol: 24 g. (0.4 mol.).
Liq. SO_2: 100 ml., Time 5 hr., at room temperature.

Bromine g.	Mol. ratio Alcohol : Bromine	Isopropyl bromide g.	Isopropyl bromide yield, %	Higher boiling fraction 42~43°C (20 mmHg)
32	1 : 0.5	8.0	16.2	5.1
48	1 : 0.75	18.5	38.0	7.0
64	1 : 1.0	19.0	38.6	7.4
77	1 : 1.2	18.8	38.1	8.3
32 (with H_2O 7.2 g.)	1 : 0.5 : (H_2O) 1.	0.5	2.0	trace

Table III. Action of chlorine and of chlorine with a catalytic amount of bromine

Alcohol: 0.4 mol., Chlorine: 0.4 mol.
Liq. SO_2: 100 ml., Time: 5 hr. at room temperature.

Alcohol	Reagent	Mole ratio of alcohol to halogen	Alkyl chloride b.p., °C	Alkyl chloride Yield, %
n-Butyl	Cl_2	1 : 1	77~79	11.6
	Cl_2+Br_2	1 : 1 : 0.02	77~79	11.6
Isobuyl	Cl_2	1 : 1	60~65	trace
	Cl_2+Br_2	1 : 1 : 0.02	60~65	2.4
sec-Butyl	Cl_2	1 : 1	67~69	6.2
	Cl_2+Br_2	1 : 1 : 0.02	67~69	3.5
tert-Butyl	Cl_2	1 : 1	49~53	18.6
	Cl_2+Br_2	1 : 1 : 0.02	49~53	17.2
Cyclohexyl	Cl_2	1 : 1	98~102	6.5
	Cl_2+Br_2	1 : 1 : 0.02	98~102	9.4

A considerable isomerization is also expected in the alkyl cation, but the writers are not yet in a position to discuss the structures of the carbon skeletons of the dimeric alkyl bromides produced.

In the bromine substitution of isopropyl alcohol, besides 38.6% of isopropyl bromide, a hexyl bromide, (probably 2-bromo-4-methylpentane) b. p. 42~43°C (20 mmHg) was formed in a yield of 21.2% of the theoretical. This substance afforded a hexyl acetate of b. p. 135~140°C with potassium acetate in glacial acetic acid. Its hydrolysis gave a hexene, C_6H_{12}, b. p. 70~73°C, possessing a comphor-like aroma, and a small amount of a hexyl alcohol, b. p. 125~130°C.

Similar phenomena were observed in the reactions of secondary or tertiary butyl alcohols with bromine, large amounts of octyl bromides being formed besides butyl bromides. Further, formation of dibromobutanes were also observed, which are assumed to have been formed by the addition of bromine to butenes produced from butyl cations.

In the alicyclic alcohols, it was observed that the formation of the cycloalkylene dibromides predominated over the cyclohexyl bromides. Thus, cyclohexyl alcohol afforded dibromocyclohexane in a yield of 64.2% and *d*-borneol afforded 71.2% of 2,10-dibromocamphane, indicating that a path through the Wagner-Meerwein rearrangement was followed.

The relation between the yield of the alkyl bromide vs. the amount of bromine was examined with isopropyl alcohol, the results being listed in Table II.

Only a small yield of the products was obtained with hydrogen bromide formed in the liquid sulfur dioxide[7] according to the formula:

$$Br_2 + 2H_2O + SO_2 \longrightarrow 2HBr + H_2SO_4 \qquad (16)$$

In the next place, the action of chlorine on aliphatic alcohols was examined. The results shown in Table III indicate that chlorine is not so reactive as bromine, the yields being depressed. Subsequently, a small amount of bromine was added to chlorine (0.02 mol. of bromine to one mol. of chlorine) in expectation of a better result as in the case of the Beckmann rearrangement by chlorine[1] in liquid sulfur dioxide, but it proved ineffective.

Experimental

Materials.—*1. Liquid Sulfur Dioxide.*—Liquid sulfur dioxide, dehydrated and purified in the plant and filled in an iron cylinder, was used. The water content measured in accordance with the principle of the Karl Fischer method

7) J. Ross, T. H. Percy, R. L. Brandt, A. I. Gebhart, J. E. Mitchell and S. Yolles, *Ind. Eng. Chem.*, 34, 924 (1942).

was 0.05～0.2%. This was further dehydrated with concentrated sulfuric acid and then with phosphorus pentoxide to bring the water content below 0.01%. However, practically no difference was found between these two kinds of liquid sulfur dioxide with respect to the reaction,

2. Alcohols. — All the alcohols were commercial products of the purest grade, which were dehydrated by the known method and redistilled.

3. Bromine. — Commercial bromine was washed with water and with concentrated sulfuric acid, dehydrated, and distilled once.

4. Chlorine.—Commercial chlorine was dehydrated with concentrated sulfuric acid and distilled.

5. Reaction Vessel.—A glass pressure bottle of 200 ml. capacity was used.

Experimental Method.—The amount of alcohols was 0.4 mol. each and that of bromine 0.4 mol. The solution of bromine in 50 ml. of liquid sulfur dioxide was added gradually to 50 ml. of liquid sulfur dioxide containing the alcohol at room temperature (20°C). The mixture was allowed to stand for a definite time. The reaction of secondary and tertiary alcohols occurred at once, the bromine color disappeared immediately, and the liquid sulfur dioxide solution became turbid. Sulfuric acid separated out as a colorless, translucent liquid on the bottom of the pressure bottle. After a definite time (usually 5 hr.), liquid sulfur dioxide was distilled off and the residual liquid was poured into ice water. The unchanged alcohols, especially butyl alcohols and lower aliphatic alcohols, dissolved out in water. The oily layer, after repeatedly being washed with water and with 1 N sodium carbonate solution and again with water, was dried over calcium chloride, and distilled to examine the amount of the lakyl halide formed.

Results.—*(A) Reaction of Alcohols with Bromine.*—1. Ethyl alcohol: From the reaction of 18.4 g. (0.4 mol.) of ethyl alcohol and 64 g. (0.4 mol.) of bromine in 100 ml. of liquid sulfur dioxide for 5 hours, 3.0 g. of a bromide was obtained. B. p. 37～38°C. Yield, 2.0 g. (5.3%). $n_D^{26}=1.425$. The bromide was identified with ethyl bromide by deriving propioanilide from it. Namely, the bromide was changed into a Grignard reagent and the latter was treated with phenyl isocyanate, and propionanilide (m. p. 104°C)[8] was obtained, which showed no depression of the melting point on admixture with an authentic specimen.

2. *n*-Propyl alcohol: A mixture of 24.0 g. (0.4 mol.) of *n*-propyl alcohol, 64 g. of bromine, and 100 ml. of liquid sulfur dioxide was kept at 20°C for 5 hr. and 3.0 g. of a crude halide was obtained. Redistillation afforded 2.5 g. (5.1%) of colorless, translucent liquid, b. p. 68～70°C, $n_D^{26}=1.435$. The bromide was identified with *n*-propyl bromide by deriving butyranilide, m. p. and mixed m. p. 91～92°C.[8]

3. Isopropyl alcohol: A mixture of 24 g. (0.4 mol.) of isopropyl alcohol, 64 g. (0.4 mol.) of bromine, and 100 ml. of liquid sulfur dioxide was kept at 20°C for 5 hr. The bromine color disappeared immediately after the starting of the reaction, the whole liquid became orange-colored and turbid, and ca. 10 ml. of a colorless, translucent liquid was deposited on the bottom. After 5 hr. this liquid was collected and titration gave a value corresponding to 13.2 g. of sulfuric acid. Liquid sulfur dioxide was then evaporated, the residual liquid was poured into ice water, and the oily layer was repeatedly washed with water and with 1 N sodium carbonate as described above. The crude halide thus obtained (30.0 g.) was fractionally distilled into the following fractions:

1) b. p. 59～60°C (760 mmHg), $n_D^{26}=1.425$, 19.0 g. (38.6%).

2) b. p. 42～43°C (20 mmHg), $n_D^{25}=1.525$, 7.0 g. (21.2%).

3) Residue, 0.3 g.

Fraction 1 was isopropyl bromide, forming isobutyranilide,[8] m. p. and mixed m. p. 101～102°C. Fraction 2 was a hexyl bromide (Found: Br, 47.9. Calcd for $C_6H_{13}Br$: Br, 48.41). Five grams of Fraction 2 was refluxed with 30 ml. of glacial acetic acid and 5 g. of potassium acetate, and 3.8 g. of a hexyl acetate, b. p. 135～140°C. (Found: CH_3CO, 30.11. Calcd. for $CH_3COOC_6H_{13}$; CH_3CO, 29.82), was obtained. Hydrolysis of this hexyl acetate by warming with ethanolic potassium hydroxide for 1 hr. afforded an olefin with camphor-like aroma, a hexene, C_6H_{12}, b. p. 70～73°C (Found: C, 85.98, H, 14.02, bromine number, 188.0. Calcd. for C_6H_{12}: C, 85.63, H, 14.37%, bromine number, 190.5).

Further, 0.3 g. of an oil, b. p. 125～130°C, assumed to be a hexyl alcohol, was obtained. This hexyl bromide may be assumed to be 2-bromo-4-methylpentane.

4. *n*-Butyl alcohol: A mixture of 30.0 g. (0.4 mol.) of *n*-butyl alcohol, 64 g. (0.4 mol.) of bromine, and 100 ml. of liquid sulfur dioxide was similarly treated and 8.0 g. of the product was fractionated, affording 5 g. (7.2%) of *n*-butyl bromide, b. p. 99～101°C, $n_D^{26}=1.440$ (Found: Br, 57.9. Calcd. for C_4H_9Br: Br, 58.33%). Valeranilide[8], m. p. and mixed m. p. 62～63°C. Further distillation of the residual

8) R. L. Shriner and R. C. Fuson, "Sstematic Identification of Organic Compounds", John Wiley & Sons, Inc., New York (1948), p. 255.

liquid afforded about 2.5 g. of an oil which boiled over a wide range of 40~90°C at 10 mmHg.

5. Isobutyl alcohol: Thirty grams of isobutyl alcohol was treated with bromine as above and 17.0 g. of the reaction product was obtained. This liquid was fractionated into three components.

1) b. p. 88~91°C, 1.5 g. (2.7%).
2) b. p. 32~38°C (10 mmHg), 9.0 g. (8.6%).
3) b. p. 90~99°C (10 mmHg), 5.0 g. (5.7%).

Fraction 1, $n_D^{26}=1.438$, was isobutyl bromide, affording isovaleranilide[8], m. p. and mixed m. p. 108~109°C. Fraction 2 was an octyl bromide, (Found: Br, 40.25, Calcd. for $C_8H_{17}Br$: Br, 41.38%). Fraction 3 corresponded to dibromo-isobutane (Found: Br, 72.56. Calcd. for $C_4H_8Br_2$: Br, 73.61%).

6. *sec*-Butyl alcohol: Thirty grams (0.4 mol.) of *sec*-butyl alcohol was similarly treated and 60.0 g. of the reaction product was obtained and fractionated.

1) b. p. 90~92°C, 18.0 g. (32.5%).
2) b. p. 41~44°C (12 mmHg), 25.0 g. (51.8%).
3) b. p. 90~110°C (10 mmHg), 2.0 g. (2.3%).
4) Residue, 3.1 g.

Fraction 1, $n_D^{26}=1.438$, was *sec*-butyl bromide (Found: Br, 57.60. Calcd. for C_4H_9Br, 58.3%), forming α-methylbutyranilide[9], m. p. and mixed m. p. 107~108°C. Fraction 2 was an octyl bromide (Found: Br, 40.59. Calcd. for $C_8H_{17}Br$; Br, 41.38%), and Fraction 3 was assumed to be a dibromobutane with some impurity (Found: Br, 70.56. Calcd. for $C_4H_8Br_2$: Br, 73.61%).

7. *tert*-Butyl alcohol: Thirty grams (0.4 mol.) of *tert*-butyl alcohol was similarly treated and 69.0 g. of the oil formed was fractions.

1) b. p. 71~73°C, 13.0 g. (23.4%).
2) b. p. 46~51°C (20 mmHg), 31.5 g. (65.1%).
3) b. p. 80~110°C (20 mmHg), 4.0 g. (4.5%).
4) Residue, 7.0 g.

Fraction 1 was *tert*-butyl bromide, $n_D^{26}=1.655$, (Found: Br, 41.96. Calcd. for C_4H_9Br: Br, 41.38%). Trimethylacetanilide, m. p. and mixed m. p. 127~128°C. Fraction 2 was an octyl-bromide, $n_D^{26}=1.655$ (Found: Br, 41.96. Calcd. for $C_8H_{17}Br$: Br, 41.38%). Ten grams of Fraction 2 was refluxed with 10 g. of potassium acetate and 30 g. of glacial acetic acid for 1 hr. and 60 g. of an octyl acetate, with characteristic sweet aroma, was obtained, b. p. 120~126°C, $n_D^{26}=1.644$ (Found: CH_3CO, 20.5. Calcd. for $C_{10}H_8O_2$: CH_3CO, 21.04%). Hydrolysis of this substance with ethanolic potassium hydroxide afforded 3 g. of a colorless, translucent oil, b. p. 118~121°C, with camphor-like odor (Found: bromine number, 139.0. Calcd. for C_8H_{16} (F_1); bromine number, 142.3). Fraction 3 was assumed to be a dibromobutane with some impurity (Found: Br, 71.74. Calcd. for $C_4H_8Br_2$: Br, 73.61%).

The amount of sulfuric acid formed during this reaction was 13.0 g. (titrated as H_2SO_4).

8. Isoamyl alcohol: A similar treatment of 35.0 g. (0.4 mol.) of isoamyl alcohol afforded 4.0 g. of a product which was distilled to give 3.0 g. (6.6%) of isoamyl bromide, b. p. 120~121°C, $n_D^{26}=1.443$. Isohexananilide[8], m. p. and mixed m. p. 107~108°C.

9. Cyclohexyl alcohol: Forty grams (0.4 mol.) of cyclohexyl alcohol was treated with 0.4 mol. of bromine in 100 ml. of liquid sulfur dioxide at room temperature for 5 hr. to give 72.2 g. of an oily product, which was fractionated into two fractions:

1) cyclohexyl bromide, b. p. 55~61°C (20 mmHg), 9.2 g. (14.1%), $n_D^3=1.496$ (Found: Br, 50.4. Calcd. for $C_6H_{11}Br$; Br, 49.0%).

2) dibromocyclohexane, b. p. 97~108°C (20 mmHg), 62.2 g. (64.2%).

The dibromocyclohexane was redistilled, b. p. 98~101°C (17 mmHg), $n_D^{24}=1.549$ (literature[9]: b. p. 100°C (13 mmHg), $n_D^5=1.5506$) (Found: Br, 65.0. Calcd. for $C_6H_{10}Br_2$: Br, 66.0%).

10. *d*-Borneol: *d*-Borneol, m. p. 201~204°C, $[\alpha]_D^{21.5}=+31.88°$, was prepared, according to the procedure of Amagasa et al., by reduction of *d*-camphor with sodium in liquid ammonia followed by rectification and recrystallization. A similar treatment of 15.5 g. (0.1 mol.) of *d*-borneol and 17.6 g. (0.11 mol.) of bromine in 140 ml. of liquid sulfur dioxide at room temperature for 18 hours afforded 26.0 g. of products, from which 11.6 g. of dark crystals were separated and recrystallized from ethyl alcohol in white crystals of m. p. 90~91°C. From the filtrate, 9.7 g. of an oil was distilled at 120~134°C (4.5 mmHg) which after recrystallization showed m.p. 90~91°C, alone and in admixture with the above specimen. Total yield, 71.2%. $[\alpha]_D^{16}=-15.0°$. This was identified with 2, 10-dibromocamphane[11] by a mixed melting point. The small amount of the forerun of the filtrate, b. p. 70~75°C (3 mmHg), 1.8 g. and b. p. 110~122°C (4 mmHg), is yet unidentified. (The optical rotations were measured in 95% ethyl alcohol in a tube of 2 cm. length.)

(B) The Mole Ratios of Bromine to Alcohols and Yields of Alkyl Bromides.—Similar treatments were carried out with various mole ratios of bromine to isopropyl and to *tert*-butyl alcohols.

(C) The Preparation of Alkyl Bromides with 1 mol. of Bromine and 2 mol. of Water in Liquid Sulfur Dioxide.—The procedures were

9) S. Winstein, *J. Am. Chem. Soc.*, **64**, 2792 (1942).
10) M. Amagasa and K. Yamaguchi, presented at the 6th Annual Meeting of the Chemical Society of Japan, April, (1953), Tokyo.
11) P. Lipp and F. Lausberg, *Ann.*, **436**, 274 (1924).

similar to those described above but two mol. of water was added to liquid sulfur dioxide to produce hydrogen bromide[7].

(D) Action of Chlorine with a Small Amount of Bromine on Alcohols in Liquid Sulfur Dioxide.—In these experiments, chlorine or chlorine with 0.1 mol. of bromine to improve the yield of the chlorides was applied to aliphatic alcohols, the procedures being essentially similar as in the case of bromine.

Summary

The reaction between aliphatic alcohols and bromine in liquid sulfur dioxide at room temperature was investigated in order to obtain further information concerning the behavior of bromine in liquid sulfur dioxide. The reactivity of the alcohols increased in the order of primary <secondary<tertiary.

In addition to alkyl bromides, higher boiling bromides were also produced. The latter resulted from further elimination, dimerization and bromine addition. The rather complicated feature of the bromination is attributable to the presence of sulfur trioxide secondarily formed in the reaction mixture.

The microanalysis was performed by Miss Yoko Endo. A part of the expenses for the present work was defrayed from the Grant in Aid for Scientific Research from the Ministry of Education. The liquid sulfur dioxide was kindly donated by the Befu Chemical Industries, Ltd. All this help is gratefully acknowledged.

Chemical Research Institute of Non-aqueous Solutions
Tohoku University
Katahiracho, Sendai

On the Oxidation of β-Aminopropionitrile. I

By Toshihisa TACHIKAWA

(Received January 18, 1960)

It is known that β-aminopropionitrile which can easily be synthesized from acrylonitrile and ammonia, turns to a viscous liquid and finally to a resinous solid, when it is stored in the air for a prolonged period. However, very little has been investigated on the polymerization of aminonitriles.

Akabori[1] proposed in his interesting theory of protein formation, that the protein-like polyamide would result from the polymerization of α-aminonitrile, $NH_2CH_2CN \rightarrow (-CH_2-CO-NH-)_x$. However he gave no experimental proof on his theory.

Whitmore and Mosher[2] have studied the polymerization of β-aminopropionitrile, and stated that the reaction proceeds according to the following scheme,

$$NH_2CH_2CH_2CN \rightarrow CH_2=CHCN+NH_3$$

$$CH_2=CHCN \rightarrow \left[-CH_2-\underset{CN}{CH}-CH_2-\underset{CN}{CH}-\right]_x$$

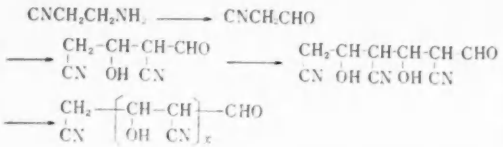
where the reaction was initiated by the decomposition of β-aminopropionitrile into acrylonitrile and ammonia, then followed by the polymerization of acrylonitrile. However, it does not seem probable that the polymerization of acrylonitrile will occur in the presence of ammonia, since ammonia is known as a polymerization inhibitor. It is also unlikely that polymerization proceeded through a viscous product because polyacrylonitrile should precipitate from the reaction mixture, as it is not soluble in acrylonitrile, water and ammonia.

As our resinous product obtained from β-aminopropionitrile was found to be easily soluble in methanol, glacial acetic acid and formic acid, this material must be quite different from the polyacrylonitrile obtained by Whitmore and his co-worker.

This paper is to report on the result of our detailed investigation on the reaction products and the mechanism of polymerization reaction of β-aminopropionitrile.

It is well known, that the aliphatic amino group is easily oxidized to aldehyde. So that β-aminopropionitrile could be oxidized in an analogous manner and the resulting aldehyde may polymerize by aldol condensation according to the following reaction scheme;

1) S. Akabori, *Kagaku*, **25**, 54, (1955).
2) F. C. Whitmore and H. S. Mosher, *J. Am. Chem. Soc.*, **66**, 725 (1944).



If this is the case, the resulting polymer may be acetylated and may give rise to the characteristic infrared absorption bands due to –OH and –CN groups. However, it was not successful to acetylate the polymer with various reagents, and no evidence for –OH or –CN was found from the infrared absorption spectrum of this compound as shown in Fig. 1. From these findings, it is believed that the polymerization reaction proceeds in a different way.

The colorless amorphous solid, obtained by treating the resinous polymer with anhydrous acetone to remove water and unreacted β-aminopropionitrile, was found to be very hygroscopic and to decompose at 120°C, resulting in the yellow mass which did not melt even above 250°C. The results of the molecular weight determination and the elementary analysis of this compound indicated the formular of $C_{12}H_{22}N_6O_5$ and the infrared absorption spectrum is shown in Fig. 1. As is clear from the spectrum, some differences were observed between the monomeric aminonitrile and the polymeric product.

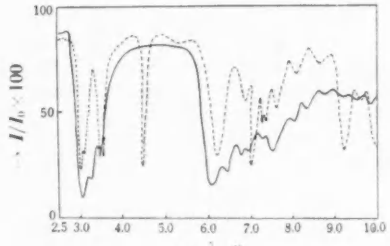


Fig. 1. Infrared absorption spectra.
---- β-Aminopropionitrile
—— Polymerization product

Two bands of N–H stretching vibration at 3.0 and 3.05 μ, splitted into three bands and shifted to 3.1, 3.3 and 3.4 μ. A band due to N–H bending vibration at 6.25 μ shifted to 6.45 μ, and a new band at 6.05 μ due to C=O stretching vibration appeared. These characteristic absorptions resembled very much those of the secondary amide such as *N*-methylacetamide which was investigated by Davis, Evans and Jones[3] as a model compound for polypeptides, and those of 6,6-Nylon. Therefore, it is very likely that this compound has a –CONH– linkage in its structure.

Regarding a new band at 6.70 μ, it is probable, that this band is due to the azoxy group. The reason for this assignment is that the position of this band coincided with that of the band arising from the stretching vibration of the azoxy group which was confirmed by Langley and others[4] in their infrared absorption study on azoxy compounds, such as 2-azoxy-2 : 5-dimethylhexane and ethyl 2-azoxybutylate.

Ultraviolet absorption spectra of these two compounds were also observed by the same authors, and their absorption maxima were reported as having two maxima at 223 and 280 mμ. The ultraviolet absorption spectrum of the polymer is also found to show two maxima at 215 and 275 mμ as shown in Fig. 2, and the locations of maxima are in good agreement with those of azoxy compounds if one considers the hypsochromic shift due to the use of water as a solvent.

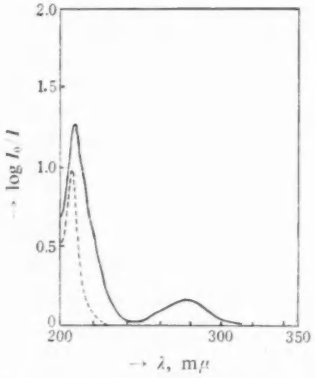


Fig. 2. Ultraviolet absorption spectra.
---- β-Aminopropionitrile
—— Polymerization product

In order to confirm the evidences of the azoxy group and –CONH– linkage in the polymer, the reduction of this compound was carried out, using concentrated hydrochloric acid and tin. The infrared absorption spectrum of the reaction product, which was obtained as colorless crystals, showed five bands at 5.85, 6.30, 6.70, 7.15 and 7.55 μ, as shown in Fig. 3, indicating the presence of amino acid, because these bands are characteristic of the hydrochloride of amino acid. And this supports the evidence of –CONH– linkage in the mother compound.

3) M. Davis, J. C. Evans and R. L. Jones., *Trans. Faraday Soc.*, **51**, 761 (1955).

4) B. W. Langley, B. Lythgoe and N. V. Riggs, *J. Chem. Soc.*, **1951**, 2309.

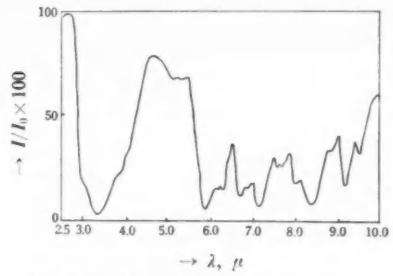


Fig. 3. Infrared absorption spectrum of reduction product of polymerization product.

Further investigation on the reduction product showed that the product was β-alanine and that this was obtained quantitatively by the reduction of the polymer. The presence of $-CONH_2$ group was also proved by the fact that ammonium chloride was found in the reduction products, and it was further confirmed that the nitrile group did not hydrolyze to -COOH group during the reaction, because no metallic salt can be precipitated from the aqueous polymer solution.

The presence of an azoxy group was also confirmed by the rapid evolution of gaseous nitrogen when the aqueous solution of the polymer was decomposed with sulfuric acid at room temperature. The fact that the aqueous solution of polymer was neutral, in contrast to the strong basicity of β-aminopropionitrile, indicated the polymer did not have a free amino group.

After considering above findings, the structure of resinous polymer obtained by standing β-aminopropionitrile in the air, is proposed as below;

$$\begin{matrix} \bar{O}\text{-}\overset{+}{N}\text{-}CH_2CH_2CONHCH_2CH_2CONH_2 \\ \quad\; \| \\ \quad\; N\text{-}CH_2CH_2CONHCH_2CH_2CONH_2 \end{matrix}$$

And the scheme of decomposition to yield β-alanine during the reduction may be shown as follows;

$$\begin{array}{lll} \overset{+OH}{O\text{-}N\text{-}CH_2CH_2CO} & \overset{+H}{NHCH_2CH_2CO} & \overset{+OH}{\overset{+H}{NH_2}} \\ \quad\; \| \overset{+H_2}{\cdots\cdots} & & \\ \quad\; N\text{-}CH_2CH_2CO & NHCH_2CH_2CO & \overset{+OH}{\overset{+H}{NH_2}} \end{array}$$

The rate of polymerization of β-aminopropionitrile was accelerated by adding small amounts of water and bubbling the air into the reaction mixture. The polymer thus obtained was found to be exactly identical to that obtained by standing the monomer in the air, and since one mole of ammonia and a small amount of hydrogen peroxide were formed from two moles of β-aminopropionitrile, the mechanism of this polymerization can be considered as the oxidation reaction accompanied by the evolution of ammonia in the presence of water as shown in Table I.

Table I. The air oxidation reaction

Nitrile g.	H_2O g.	Gas used	Gas blown l.	Product g.	NH_3 generated g.
15	0	Air	2800	No change	None
15	1.0	Air	560	White solid 15.0	1.80
15	3.8	Air	560	White solid 15.8	1.91
15	6.0	Air	720	White solid 15.1	1.87
15	1.0	N_2	2020	No change	None

Experimentals

1. Raw Material and Sample.—β-Aminopropionitrile was synthesized from acrylonitrile and aqueous ammonia according to Buc, Ford and Wise[5], and the fraction boiling at 86.3~87.0°C/21 mmHg was collected.

A polymeric sample was prepared by standing β-aminopropionitrile for two months in the air, followed by washing the resulting solid mass with anhydrous acetone to remove unreacted aminonitrile and water. A white amorphous and very hygroscopic solid was thus obtained, and dried in a vacuum. This sample decomposed at 120°C, yielding a product which did not melt above 250°C.

Found: C, 43.27; H, 6.69; N, 25.04%; mol. wt., detd. by freezing point depression using acetic acid. 311.1. Calcd. for $C_{12}H_{22}N_6O_5$: C, 43.63; H, 6.71; N; 24.22%; mol. wt., 330.4.

Ultraviolet absorption spectra were measured for the aqueous solution of 0.0651 g./l. A Shimadzu Model QR 51 spectrophotometer with 1 cm. quart celles was employed. Infrared absorption spectra were recorded with a Perkin Elmer Model 21 spectrophotometer on the sample film which was prepared by evaporating 2—3 drops of 3% methanolic solution on the sodium chloride window.

2. Reduction of Polymer.—Twelve grams sample was dissolved into 15 g. of water, and the solution was treated with 19.1 g. of metallic tin and 75 g. of 35% hydrochloric acid at 95~100°C for 10 hr. under stirring until the tin dissolved completely.

The resulting solution was diluted with four volumes of water and stannous ion was precipitated with hydrogen sulfide. The filtrate was concentrated under reduced pressure to get a crystalline product I which was found to be ammonium chloride.

After filtering it, the filtrate was subjected to further evaporation until one could get 16.5 g. of the second crystalline product II. The infrared absorption spectrum was measured on this material II using the potassium bromide pellet method.

The product II was then dissolved in a small amount of water and after neutralizing the solution with 1 N sodium hydroxide solution, it was boiled

5) S. R. Buc, J. H. Ford and E. C. Wise, *J. Am. Chem. Soc.*, 67, 92 (1945).

with an excess amount of freshly prepared cupric hydroxide. The hot filtrate was concentrated and cooled to get dark blue crystals. This was re-dissolved in water, and cupric ion was precipitated with hydrogen sulfide. When the filtrate was evaporated almost to dryness, a crystalline product III was obtained.

Found: C, 40.27; H, 7.86; N, 15.81. Calcd. for $C_3H_7NO_2$: C, 40.44; H, 7.92; N, 15.72%.

3. Decomposition with Sulfuric Acid.—According to Curtius procedure, 0.0818 g. of the sample was dissolved in 1.0 ml. of water and 1.5 ml. of 98% sulfuric acid was dropped into this solution with gentle heating. The evolving gas, collected over a 10% potassium hydroxide solution, measured 5.34 ml. As this gas was not absorbed by the alkaline pyrogallol solution, ammoniacal cuprous chloride solution or barium hydroxide solution, it was found to be pure nitrogen. The amount of nitrogen evolved was 8.17% per mole of the sample.

4. Determination of Neutralization Equivalent.—A solution of 0.0999 g. sample in 50 ml. of water was titrated with 0.1 N hydrochloric acid using methyl orange as an indicator. Neutralization equivalent was calculated as 951.4 from the titre of 1.05 ml. This value is considerably larger than that of β-aminopropionitrile (69.5), and the sample can be considered as almost neutral material.

5. Formation of Metallic Salts.—No precipitation was observed when barium hydroxide, cupric nitrate or silver nitrate solution was added to the aqueous sample solution.

6. Air Oxidation.—Decarbonated air or nitrogen was bubbled into β-aminopropionitrile in a 100 ml. reaction flask at the rate of 20 l./hr. The increase of viscosity and the rapid evolution of gas were observed during she introduction of air, which was continued until the content rapidly increased its viscosity, resulting in a resinous material. This product was treated with anhydrous acetone to remove unreacted β-aminopropionitrile and water, then it was dried under reduced pressure. The product was obtained as a very hygroscopic white amorphous mass.

Amount of ammonia evolved during the reaction, was determined by passing the gas into sulfuric acid, and the presence of hydrogen peroxide in the gas was detected by passing it into barium hydroxide solution, followed by the dissolution of the resulting white precipitate into hydrochloric acid, which turned to red with an addition of potassium iodide.

The author wishes to express his sincere thanks to Professor Ryozo Goto of Kyoto University for his kind advice, and also to Mr. Murataro Nakashima, President of Kunoshima Chemical Co., for his permission to publish this paper.

Research Laboratory of Kunoshima Chemical Industry Co., Ltd.
En-ichi-cho, Mihara

The Self-diffusion of Surface Active Agents in Nylon

By Makoto HAYASHI

(Received January 20, 1960)

Adsorption of surface active agents on various solid surfaces such as metal, carbon, wool, cotton, acetate and polyamide fibers has been studied by many workers[1-5] as the fundamental contributions to the study of the physicochemical properties of detergents or as the first step to study the mechanism of detergency. However, the process of adsorption of detergents by the solid is rather complicated and there are a number of problems to be solved by further investigation.

We have studied[6] the adsorption of the surface active agents such as sodium dodecyl sulfate, sodium tetradecyl sulfate as well as inorganic ions such as sulfate, chloride, phosphate and calcium ions on the surface of nylon plate. The quantities of these substances taken up by nylon are very great and a very long time is required to reach an adsorption equilibrium. From these facts, it seems that the penetration of the adsorbed substances into the solid body plays an important role in the process of sorption and so the measurement of diffusion in solid is considered to be significant in a study of the adsorption of these substances to polymer solid.

In this paper, the self-diffusion rates in nylon

1) A. L. Meader and B. A. Fries, *Ind. Eng. Chem.*, **44**, 1636 (1952).
2) L. Hsiao and H. N. Dunning, *J. Phys. Chem.*, **59**, 362 (1955).
3) A. Fava and H. Eyring, ibid., **60**, 890 (1956).
4) H. C. Evans, *J. Colloid Sci.*, **13**, 537 (1958).
5) R. D. Vold and N. H. Sivaramakrishnan, *J. Phys. Chem.*, **62**, 984 (1958).
6) M. Hayashi and Y. Nakazawa, *Radioisotopes*, **8**, 149 (1959).

of sodium dodecyl sulfate and sodium tetradecyl sulfate, and also of sulfuric acid and hydrochloric acid were studied, using the nylon plates and the radioactive tracers.

Experimental Method

The measurement was carried out by the surface radioactivity decrease method, which consists in measuring the decrease in radioactivity of the surface with the time due to the diffusion of the radioactive tracer into nylon. This method has the following advantages: the measurement is accomplished in a short diffusion time when the suitable β-ray energy such as that of ^{35}S is used and it is not necessary to use any particular arrangements.

A nylon plate was 2.5 mm. in thickness, 2.5 cm. $\times$2.5 cm. in size. It was cleaned by washing with ether, acetone and hot water before use.

The surface active agents used in this experiment were sodium dodecyl surfate (SDS) and sodium tetradecyl sulfate (STS). They were prepared from pure dodecyl alcohol and tetradecyl alcohol, respectively, and their purity was confirmed surface-chemically. The radioactive SDS and STS, both tagged with ^{35}S, were prepared from corresponding alcohols and radioactive sulfuric acid ($H_2{}^{35}SO_4$). Radioactive sulfuric acid ($H_2{}^{35}SO_4$) and hydrochloric acid ($H^{36}Cl$) were used for the measurement of diffusion of the sulfate ion and the chloride ion, respectively. The absorption coefficients of the β-rays of ^{35}S and ^{36}Cl in nylon were estimated using the sufficiently swollen nylon film (25μ in thickness), and the values, 321 cm^{-1} for ^{35}S and 12.7 cm^{-1} for ^{36}Cl were obtained.

Diffusion experiments were accomplished as follows: the nylon plates are immersed in various concentrations of each non-radioactive solution at 33°C for about two weeks in order that the nylon might sorb the solute by a complete equilibrium amount corresponding to each concentration. (It has been shown that the sorption equilibrium is accomplished completely under these circumstances.) They are taken up from the solution, blotted by filter paper, then immersed instantly in each radioactive solution of the same concentrations as the non-radioactive solution used, for the purpose of making a surface radioactive layer, probably, by surface exchange. They are stirred sufficiently for about 0.5~2 min., then the nylon plates are taken up, washed with a large amount of water for a few seconds, blotted with filter paper and counted instantly under the G-M tube. They are then fitted in a closed glass vessel of saturated humidity which is then kept in a thermostat. When the diffusion has progressed appropriately (about 20 hr. for SDS and STS, 4 hr. for H_2SO_4, 48 hr. for HCl at 33°C, respectively), the nylon plates are removed from the vessel and their radioactivity is counted in the same manner as before, and the ratio A/A_0 (A_0: initial counts of the surface of nylon, A: its counts at time t) is estimated. A/A_0 is given by the following equation[7,8]

$$A/A_0=[1-\mathrm{erf}\{(\mu^2Dt)^{1/2}\}]\exp(\mu^2Dt) \quad (1)$$

where D is the self-diffusion coefficient of the diffusing ion, and μ is the absorption coefficient of the β-ray in nylon. A/A_0 can be plotted as a function of μ^2Dt and by using this graph D can be obtained from the measured value of A/A_0 since μ and t are known.

The quantity of each solute sorbed by nylon was measured by the same method as in the previous paper[6] and expressed as millimoles per kilogram of nylon, which was calculated by calibrating the decrease of counts owing to the absorption of radiation by nylon.

Radioactivity measurements carried out by means of a thin mica window G-M tube (1.4 mg./cm^2) and the conventional scaler. For the purpose of establishing the counting area and avoiding the edge effect on counting, a brass ring 1 mm. think and 22 mm. in inner diameter was placed on the plate when it was counted.

The change of the efficiency of the counting equipment was always checked by a radioactive standard prepared on a nylon plate. Counting times were made so long that the standard counting error was less than 1%.

Results

To confirm the compatibility of the above diffusion equation (1) in these measurements the values of A/A_0 at various diffusing time were measured at 33°C with SDS and H_2SO_4 (Figs. 1 and 2). In these figures, the curves of solid line show the calculated values obtained by the above equation and the value of D at $A/A_0=0.5$. The points are the measured ones. It is seen that this equation can be applied to calculate the diffusion coefficient in these systems.

The sorption isotherms of each substance by nylon plate were measured by the radiotracers to estimate the quantities sorbed in nylon.

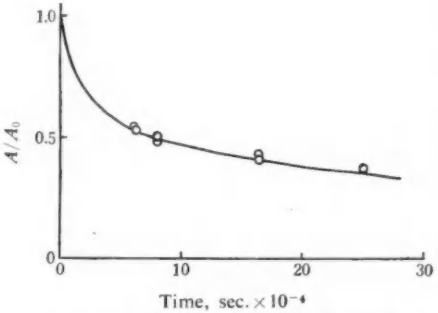


Fig. 1. The check of compatibility of the Eq. 1 by the diffusion of SDS at the concentration of 34 millimol./kg. in nylon at 33°C. The solid line is the calculated values and the points are the measured ones.

7) V. Linnenbom, M. Tetenbaum and C. Cheek, *J. Appl. Phys.*, **26**, 932 (1955).
8) F. Bueche, W. M. Cashin and P. Debye, *J. Chem. Phys.*, **20**, 1956 (1952).

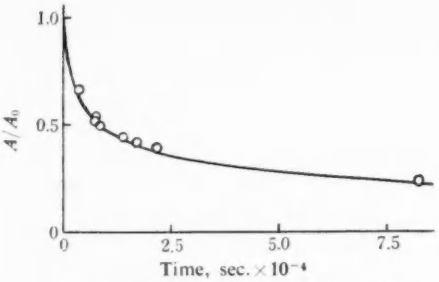


Fig. 2. The check of compatibility of the Eq. 1 by the diffusion of H_2SO_4 at the concentration of 63 millimol./kg. in nylon at 33°C. The solid line is the calculated values and the points are the measured ones.

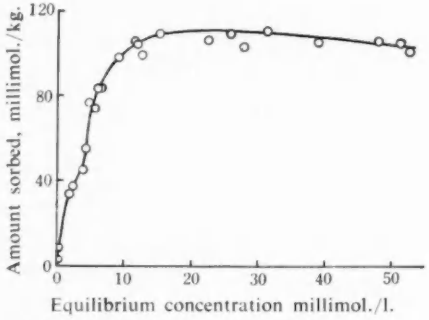


Fig. 3. The isotherm of the sorption of SDS by nylon plate at 33°C.

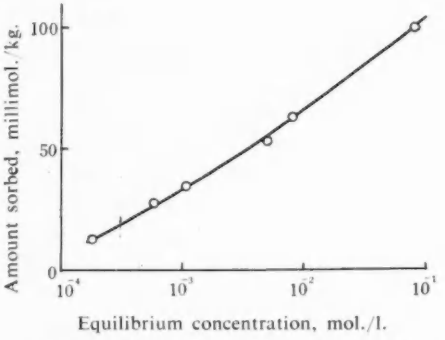


Fig. 4. The isotherm of the sorption of H_2SO_4 by nylon plate at 33°C.

The results for SDS and H_2SO_4 as examples are shown in Figs. 3 and 4, respectively.

The diffusion coefficients of SDS, STS, H_2SO_4 and HCl were measured at various concentrations of every solute in nylon. The results are shown in Figs. 5 and 6 and in Table I.

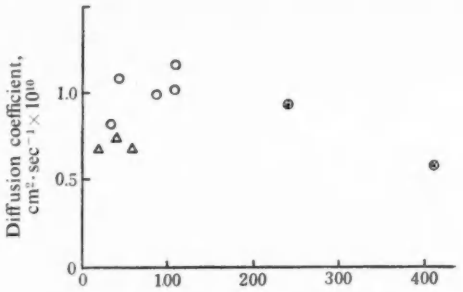


Fig. 5. Diffusion coefficient of SDS and STS at 33°C.

○ SDS
△ STS
⊙ The values of SDS obtained by the nylon plate which contained the large quantity of SDS by the lowering of the medium pH with HCl and KCl.

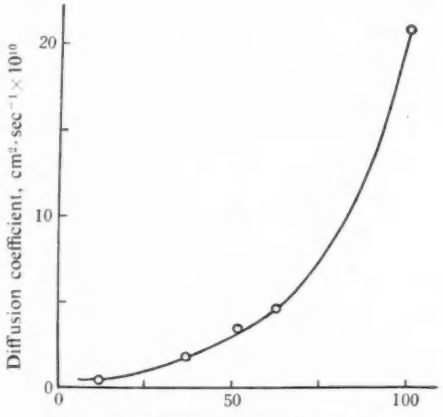


Fig. 6. Diffusion coefficient of H_2SO_4 at 33°C.

TABLE I. SELF-DIFFUSION COEFFICIENTS OF HCl AT 33°C

Concentration of HCl* mol./l.	Diffusion coefficient $cm^2 \cdot sec^{-1} \times 10^9$
10^{-1}	6.46
10^{-2}	0.120
10^{-3}	0.176

* The concentration of the solution described here is that which is in sorption equilibrium.

The diffusion coefficients of SDS and H_2SO_4 at various temperatures were measured to observe the effect of temperature on the diffusion and are shown in Table II. When $\log D$ is plotted against $1/T$, the straight lines are obtained as shown in Fig. 7, from which the

TABLE II. TEMPERATURE DEPENDENCE OF DIFFUSION COEFFICIENTS

Diffusing substance	Concentration in nylon (millimol./kg)	Diffusion coefficient $cm^2 \cdot sec^{-1} \times 10^{10}$					Activation energy, E cal./mol.
		−4°C	5°C	10°C	19°C	33°C	
SDS	102	0.00012	0.0053	0.0177	0.182	1.10	33400
H_2SO_4	63	0.0209	0.102	0.241	0.927	4.69	23200

activation energy E in the equation $D = D_0 \exp(-E/RT)$ was estimated and shown in the last column of Table II.

If the nylon sample is dried before or after the dipping in radioactive solution, the diffusion is little observed by this method at the room temperature.

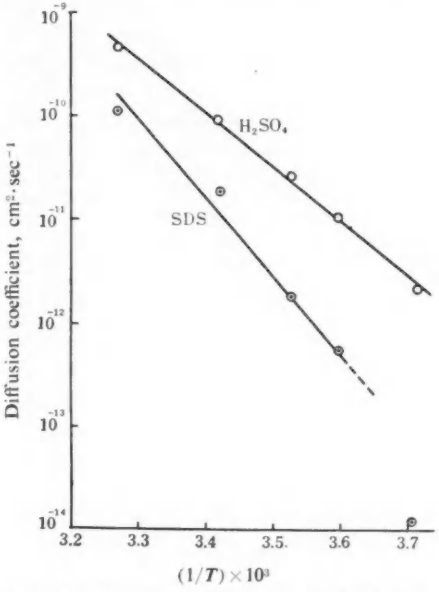


Fig. 7. Temperature dependence of diffusion coefficients of SDS (102 millimol./kg.) and H_2SO_4 (63 millimol./kg.).

Discussion

In this experiment the measurements of the self-diffusion rate were always carried out using sufficiently swollen nylon plates and it has been confirmed by weight that the swelling of each sample under the condition of the present experiment is constant throughout the whole diffusion process. The self-diffusion of the surface active agents or inorganic acids became remarkably slower as a nylon plate was dried and could not be observed at the room temperature in the completely dried state of nylon. It is evident, therefore, that these substances diffuse with the aid of water in nylon.

It is clear from the results of the present experiment that in the sorption of surface active agents and also of inorganic acids from their aqueous solutions by nylon, not only the surface phenomenon but also the diffusion into the solid body is responsible for its process towards equilibrium.

The self-diffusion coefficients of SDS and STS in nylon at 33°C are dependent, though not remarkably, on the concentration of each detergent in nylon as shown in Fig. 5. The values are comparable with those of several dyes reported by some workers[9,10], but they are much smaller than the values[11] in aqueous solutions. We suppose that the internal body of nylon is a gel of extremely close-packed network structures, water being included and forming fine canals among them. The molecules of detergent will penetrate along these canals into the internal structures of nylon, but they are immobilized by the adsorption sites to a large extent and so the rate of diffusion is reduced considerably compared with the free diffusion in aqueous systems. It is considered that the jumping of the adsorbed molecules from site to site occurs just as in the diffusion of ions in their crystals. Indeed the activation energies for diffusion of SDS and sulfuric acid are 33.4 and 23.2 kcal. per mole respectively as shown in Table II. These values are much greater than the values of many substances in aqueous systems, rather resembling these in solid crystals[12].

We suppose that the interaction of ionic detergents and nylon takes place not only by ionic force but also by van der Waals force, since the amount of detergents sorbed by nylon is much larger than that of simple inorganic ions which are considered to be bound to nylon mainly by ionic interaction[6]. This is confirmed by the fact that the amount of detergents taken up is also much larger than the value corresponding to the amine content of nylon[13,14]. The sorption of sulfuric acid

9) T. Vickerstaff, "The Physical Chemistry of Dyeing", 2nd ed. Olyver and Boyd, London (1954), pp. 152, 449.
10) M. Yoshida and H. Morita, *Sen-i-gakkaishi*, **12**, 662 (1956).
11) D. Stigter, R. J. Williams and K. J. Mysels, *J. Phys. Chem.*, **59**, 330 (1955).
12) G. E. Boyd and B. A. Soldano, *J. Am. Chem. Soc.*, **75**, 6091 (1954).
13) J. E. Waltz and G. B. Taylor, *Anal. Chem.*, **19**, 448 (1947).
14) F. C. McGrew and A. K. Schneider, *J. Am. Chem. Soc.*, **72**, 2547 (1950).

and hydrochloric acid will be caused mainly by the attachment of these substances to the amino groups or to the amide groups in the molecule of nylon. From this point of view, therefore, it is supposed that there is also a difference between detergent and inorganic acid in the mechanism of diffusion. Indeed as shown in Figs. 5 and 6, the diffusion coefficient of the detergent does not change markedly with concentration of the detergent in nylon, while that of the inorganic acid increases remarkably with its concentration. At present, however, it is difficult to give a precise mechanism for the diffusion of the detergent and inorganic acid in nylon. This will be discussed after further investigation has been performed.

Summary

The self-diffusion rates of sodium dodecyl sulfate, sodium tetradecyl sulfate, sulfuric acid and hydrochloric acid in nylon containing various amounts of each solute and their temperature dependence were measured by the surface radioactivity decrease method using ^{35}S and ^{36}Cl.

From the experimental results, it has become clear that the diffusion of the solute in nylon is important in the process towards the attainment of sorption equilibrium of these substances by nylon, and in this case the presence of water in nylon is necessary for the diffusion to take place appreciably. The diffusion coefficients were much smaller than those in their aqueous solutions.

The diffusion rates of the surface active agents depended little on the concentration in nylon compared with those of the inorganic acids. The diffusion rates of sulfuric acid and hydrochloric acid depended largely on their concentration in nylon. This will be due to a difference of the mechanism of diffusion between the former and the latter.

The activation energies of SDS and sulfuric acid for diffusion were considerably great and approximated to those of many substances in solid.

The author wishes to express his gratitude to Professor R. Matuura of Kyushu University for his kind guidance and for supplying the radioisotopic samples, and also to Toyo Rayon Co. for its supply of nylon samples.

Department of Pharmacology
Nagasaki University
School of Medicine, Nagasaki

Estimation of the Energy of Hydrogen Bonds Formed in Crystals. III. Amides

By Ariyuki AIHARA

(Received January 20, 1960)

A study of the NH···O bonding bears a particular significance in relation to the basic structure and behavior of proteins, and a fairly large amount of data of the bond have been accumulated up to now. Nevertheless, it cannot be said that the strength of the NH···O bonding has been fully examined so as to allow us to decide definitely whether this type of hydrogen bond is weaker than the OH···O bonding or not, especially in the crystalline state. Thus we find a considerable divergence in the data of the energy of NH···O bondings already published by several investigators. For instance, Coulson cites, in his well-known publication[1], 2~3 kcal./mol. for the energy of the NH···O bond, while the following data are available in the literature: 3.0 kcal./mol. (acetamide, *N*-methylacetamide)[2], (diacetylhydrazine)[3], 3.5 kcal./mol. (urea)[4], 4.2 kcal./mol. (oxamide)[5], 4.7 kcal./mol. (sulfamide)[3], 5.3 kcal./mol. (diketopiperazine)[6], ~6.0 kcal./mol.

* Present address: University of Electro-Communications, Chofu-shi, Tokyo.

1) C. A. Coulson, "Valence", Clarendon Press, Oxford (1952). p. 301.
2) A. Aihara, *J. Chem. Soc. Japan, Pure Chem. Sec. (Nippon Kagaku Zasshi)*, **73**, 855 (1952).
3) S. Takagi, R. Shintani, H. Chihara and S. Seki, This Bulletin, **32**, 137 (1959).
4) K. Suzuki, S. Onishi, T. Koide and S. Seki, ibid., **29**, 127 (1956).
5) R. S. Bradley and A. D. Care, *J. Chem. Soc.*, **1953**, 1681.
6) S. Seki, K. Suzuki and T. Koide, *J. Chem. Soc. Japan, Pure Chem. Sec. (Nippon Kagaku Zasshi)*, **77**, 346 (1956).

(δ-valerolactam and ε-caprolactam)[7], 6.1 kcal./mol. (diformylhydrazine)[4].

In this connection, the author has tried to get more knowledge on the strength of the NH···O bond, and has estimated the energy of the bond with thirteen amides in the crystalline state, following the procedure proposed in previous papers[8]. The obtained values of the energy range from 3.5 to 7.0 kcal./mol., showing that the NH···O bond is not always weaker than the OH···O bond as has hitherto been presumed.

Experimental

The measurement of the sublimation pressures was done with the same apparatus as that used before[8]. Of the materials dealt with in this investigation, formanilide and *N*-methylbenzamide were prepared in this laboratory following the usual procedure (the former from aniline and formaldehyde, the latter from methylamine and benzoylchloride). δ-Valerolactam and ε-caprolactam were kindly sent by Dr. Tsuboi, the University of Tokyo The result of the measurement of their sublimation pressures was reported before by the author[7], but, for the sake of comparison, it will be included here. 1-Phenyl-3pyrazolidone (phenydone) was sent by Dr. Kurosaki, Fuji Photo Film Co., Ltd., to whom author's thanks are due. Others were all commercial products and were purified by repeated fractional sublimation under high vacuum pressure before use.

Results and Discussion

Sublimation Pressures.—The results of measurements of sublimation pressure of thirteen amides are shown in the following tables (from Table I to Table XIII) together with their

Table I. Sublimation pressure of formanilide

T°K	P(mmHg)$\times 10^4$	T°K	P(mmHg)$\times 10^4$
297.7	6.86	311.9	28.7
301.6	10.3	313.6	34.2
308.4	20.3	314.6	37.2
309.3	22.0	316.8	45.9
310.7	25.9	317.8	49.4

$\log P(\text{mmHg}) = 10.511 - 4071.0/T$: m. p. 48.7°C.

Table II. Sublimation pressure of acetanilide

T°K	P(mmHg)$\times 10^4$	T°K	P(mmHg)$\times 10^4$
317.7	8.03	327.7	21.2
320.0	10.2	330.1	26.8
322.7	12.9	333.2	36.2
325.4	16.4	335.8	48.1

$\log P(\text{mmHg}) = 11.166 - 4533.6/T$: m. p. 115.0°C.

Table III. Sublimation pressure of *o*-acetoluidine

T°K	P(mmHg)$\times 10^4$	T°K	P(mmHg)$\times 10^4$
315.2	4.20	330.8	23.6
318.5	6.31	334.2	35.3
321.2	8.80	337.5	48.6
324.2	11.8	340.0	65.6
327.3	16.5		

$\log P(\text{mmHg}) = 12.717 - 5070.8/T$: m. p. 111.5°C.

Table IV. Sublimation pressure of *p*-acetotoluidine

T°K	P(mmHg)$\times 10^4$	T°K	P(mmHg)$\times 10^4$
330.7	6.20	337.9	13.0
332.1	7.23	340.6	17.5
333.1	7.93	343.4	23.2
334.0	8.98	346.5	31.6
335.2	10.1	350.1	46.6

$\log P(\text{mmHg}) = 12.437 - 5172.9/T$: m. p. 149°C.

Table V. Sublimation pressure of *p*-acetaminobenzaldehyde

T°K	P(mmHg)$\times 10^4$	T°K	P(mmHg)$\times 10^4$
328.3	5.59	337.4	14.9
330.1	6.73	339.9	19.4
332.6	8.82	342.4	25.2
335.1	11.5	345.4	32.8

$\log P(\text{mmHg}) = 12.516 - 5177.7/T$: m. p. 155.5°C.

Table VI. Sublimation pressure of benzanilide

T°K	P(mmHg)$\times 10^4$	T°K	P(mmHg)$\times 10^4$
353.0	3.93	363.7	10.6
355.4	4.73	366.7	13.7
357.0	5.64	368.9	16.4
360.2	7.53		

$\log P(\text{mmHg}) = 11.264 - 5181.2/T$: m. p. 163.5°C.

Table VII. Sublimation pressure of benzamide

T°K	P(mmHg)$\times 10^3$	T°K	P(mmHg)$\times 10^3$
325.3	1.04	334.3	2.67
327.4	1.31	335.4	3.00
328.4	1.39	337.0	3.49
330.0	1.70	338.9	4.37
332.9	2.33	341.9	5.86

$\log P(\text{mmHg}) = 12.579 - 5063.6/T$: m. p. 130.5°C.

Table VIII. Sublimation pressure of *N*-methylbenzamide

T°K	P(mmHg)$\times 10^4$	T°K	P(mmHg)$\times 10^4$
308.0	7.06	320.9	27.3
310.6	9.16	323.7	35.3
313.0	11.9	326.8	47.8
315.7	15.7	329.3	60.9
318.5	21.0		

$\log P(\text{mmHg}) = 11.358 - 4469.8/T$: m. p. 79.8°C.

7) A. Aihara, ibid., **74**, 631, 634 (1953).
8) A. Aihara, This Bulletin, **32**, 1242 (1959); **33**, 194 (1960).

TABLE IX. SUBLIMATION PRESSURE OF α-PHENYLACETAMIDE

T°K	P(mmHg)×10⁴	T°K	P(mmHg)×10⁴
330.0	5.18	340.9	16.4
332.9	7.12	344.0	21.9
335.2	9.09	348.0	31.7
338.2	12.4	351.8	46.5

$\log P(\text{mmHg}) = 12.242 - 5123.7/T$: m. p. 158.5°C.

TABLE X. SUBLIMATION PRESSURE OF N-ACETYL-α-NAPHTHYLAMINE

T°K	P(mmHg)×10⁴	T°K	P(mmHg)×10⁴
337.9	2.66	348.0	6.93
341.0	3.56	350.6	8.80
342.3	4.18	354.5	12.2
345.5	5.47	359.7	20.8

$\log P(\text{mmHg}) = 10.959 - 4912.2/T$: m. p. 159°C.

TABLE XI. SUBLIMATION PRESSURE OF 1-PHENYL-3-PYRAZOLIDONE

T°K	P(mmHg)×10⁴	T°K	P(mmHg)×10⁴
327.8	4.18	340.4	13.2
331.0	5.46	343.9	17.5
333.9	6.85	348.2	24.6
336.8	9.09		

$\log P(\text{mmHg}) = 10.025 - 4397.5/T$: m. p. 120.8°C.

TABLE XII. SUBLIMATION PRESSURE OF δ-VALEROLACTAM

T°K	P(mmHg)×10³	T°K	P(mmHg)×10³
293.9	2.81	304.7	8.35
294.7	3.09	306.7	10.35
297.0	3.96	308.4	11.66
298.9	4.67	310.6	13.84
300.7	5.65	312.5	15.70
302.5	6.63		

$\log P(\text{mmHg}) = 10.692 - 3891.0/T$: m. p. 39.3°C.

TABLE XIII. SUBLIMATION PRESSURE OF ε-CAPROLACTAM

T°K	P(mmHg)×10³	T°K	P(mmHg)×10³
294.0	1.18	305.2	4.21
295.2	1.38	306.9	5.03
297.1	1.72	308.7	6.02
298.3	1.93	310.6	7.43
300.4	2.44	312.2	8.70
301.9	2.92	313.9	10.03
303.7	3.56		

$\log P(\text{mmHg}) = 11.839 - 4339.8/T$: m. p. 69.8°C.

vapor pressure equations obtained by the method of least squares. The linear relation between $\log p$ and $1/T$ is found to be quite satisfactory for all the measurements here mentioned.

In addition to the sublimation pressure of the amides listed above, that of *p*-aminoacetophenone, which should be classified as an amine in the usual sense, is also shown in Table XIV. Since the molecular structure of *p*-aminoacetophenone bears some resemblance to that of acetamide, except that the amino group is separated from the carbonyl group by a phenyl nucleus, it might be expected to form NH···O bonds with neighboring molecules in the crystal.

TABLE XIV. SUBLIMATION PRESSURE OF p-AMINOACETOPHENONE

T°K	P(mmHg)×10⁴	T°K	P(mmHg)×10⁴
314.2	3.32	326.0	11.9
315.6	3.78	329.2	16.5
318.6	5.47	332.3	22.5
320.1	6.42	335.2	30.2
323.8	9.47	338.2	41.7

$\log P(\text{mmHg}) = 11.900 - 4832.6/T$: m. p. 106.2°C.

The values of the heat of sublimation (ΔH), the change of free energy at 298.2°K ($\Delta G^{298.2°}$) and of entropy (ΔS) on sublimation are shown in Table XV.

Energy of Hydrogen Bond.—Table XV shows the estimated energies (E) of hydrogen bonds (NH···O) in the fourth column, together with the hypothetical lattice energies ($\Delta H'$) obtained conventionally, assuming the additivity rule of the lattice energy, without taking into account the formation of hydrogen bonds in the crystalline state[8]. The values of E were obtained by subtracting $\Delta H'$ from ΔH.

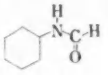
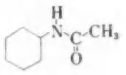
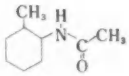
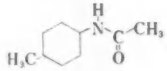
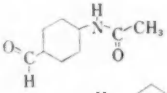
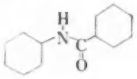
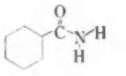
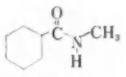
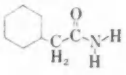
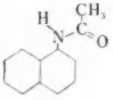
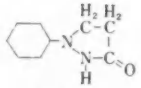
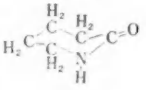
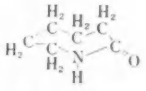
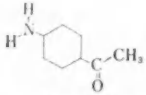
Because the crystal structures of the majority of compounds listed in Table XV are not analyzed as yet except for those of acetanilide[9] and benzamide[10], the formation of the NH···O bond is implicitly assumed for the rest of the compounds. considering the similarity of their molecular structures, this assumption would not be unreasonable.

It is clear, from the values of E in the table, that the compounds dealt with in this investigation can be classified into three categories: the first showing little or no tendency to associate mutually by NH···O bondings (benzanilide, *N*-acetyl-α-naphthylamine and 1-phenyl-3-pyrazolidone); the second forming NH···O bondings of the order of 4±1 kcal./mol. (formanilide, acetanilide, *o*- and *p*-acetotoluidine, *p*-acetaminobenzaldehyde, benzamide, *N*-methylbenzamide and α-phenylacetamide), and the third capable of making hydrogen bondings as strong as 6~7 kcal./mol. (δ-valerolactam and ε-caprolactam).

Let us first examine those which belong to category I. An alternative nomenclature of

9) C. J. Brown and D. E. Corbridge, *Nature*, **162**, 72 (1948); *Acta Cryst.*, **7**, 711 (1954).
10) B. R. Penfold and J. C. B. White, ibid., **12**, 130 (1959).

TABLE XV. VALUES OF THERMODYNAMIC FUNCTIONS AND ENERGIES OF HYDROGEN BOND

Compounds	ΔH kcal./mol.	$\Delta H'$ kcal./mol.	E kcal./mol.	$\Delta G^{298.2°}$ kcal./mol.	ΔS cal./deg. mol.
	18.63±0.11	14.5	4.1	8.22±0.21	34.9±0.3
	20.75±0.37	16.5	4.3	9.44±0.70	37.9±1.1
	23.20±0.22	18.5	4.7	9.78±0.40	45.0±0.6
	23.67±0.23	18.5	5.2	10.63±0.43	43.7±0.7
	23.69±0.23	20.0	3.7	10.54±0.44	44.1±0.7
	23.71±0.29	23.5	0.2	12.27±0.53	38.4±0.8
	23.17±0.27	14.5	8.7×(1/2)	9.94±0.51	44.4±0.8
	20.45±0.08	16.5	4.0	8.85±0.15	38.9±0.2
	23.45±0.21	16.5	7.0×(1/2)	10.67±0.46	42.8±0.8
	22.48±0.30	23.0	−0.5	11.46±0.86	37.0±0.6
	20.12±0.40	20.5	−0.4	10.37±0.75	32.7±1.2
	17.81±0.12	11.5*	6.3	7.15±0.16	35.7±0.4
	19.86±0.11	13.0*	6.9	7.64±0.22	41.0±0.4
	22.11±0.13	16.5	5.6×(1/2)	9.80±0.25	41.3±0.4

* In these cases, the increment of the lattice energy due to a CH_2 group was taken as 1.5 kcal./mol., taking into consideration the volume effect due to ring formation: the value was estimated from the heat of sublimation of cyclohexane, ~8.5 kcal./mol. (the sum of the heats of vaporization and fusion).

benzanilide is *N*-phenylbenzamide and the molecule consists of two large phenyl groups attached to both sides of a peptide linkage. These phenyl groups will inevitablly prevent the molecules from approaching each other closely to make association by the NH···O bonding. The infrared absorption data[11] prove also that the NH···O bond in the crystal of benzanilide is rather weak, if it is formed to any extent: the wavelength of the N–H absorption band in the crystalline state is 2.99 μ compared with 3.08 μ in the case of acetanilide, which is known to form NH···O bonds in the crystals[9]. *N*-acetyl-α-naphthylamine is also known to show little tendency to form hydrogen bonds with dioxane and cyclohexanone as revealed from the study of ultraviolet absorption of this substance by Baba and Suzuki[12]. The small negative value of E for this compound implies that there would be no hydrogen bond formed in the crystalline state in accordance with the result of ultraviolet absorption. This may be due to the blocking effect of a phenyl nucleus located very close to the N–H group at the α-position.

As to 1-phenyl-3-pyrazolidone, Kurosaki[13] has observed the formation of associated dimers in its CCl_4 and $CHCl_3$ solutions by the infrared and dielectric technique. But, at the same time, he noticed that its tendency to associate was much decreased in $CHCl_3$ solution due to the high polarity of the solvent which interacts with the solute molecules. Therefore, if an extrapolation is allowed to be made to the extreme case where the solute is surrounded completely by the same substance as a solvent (that is, a pure liquid or solid), we may assume that the molecule should exist as a monomer in the crystalline state without making any association, because the dipole moment of 1-phenyl-3-pyrazolidone 3.5~3.6 D[13] is much greater than that of chloroform 1.05 D[14]. The result obtained by the author shown in Table XV confirms this conclusion.

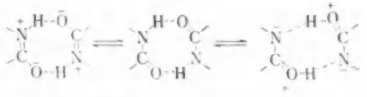
Those in category II may be regarded as forming "normal" hydrogen bonds of NH···O type, the energy of the bond being about 4 kcal. mol. The small differences between the values of these compounds might have some meaning concerning the strength of the bond formed, but, considering the approximate nature of the method of estimation, a further discussion on the difference of the strength of the bond cannot be made. It is interesting, however, to see that the methyl groups substituted to the phenyl nucleus seem to incsease the strength of the NH···O bond in the case of *o*- and *p*-acetotoluidine, while the formyl group in the para position of *p*-acetaminobenzaldehyde tends to decrease the NH···O bond-forming-power of the molecule: the weakly electron-releasing inductive and hyperconjugative effects of the methyl group might cause an increase of the negative charge on the oxygen atom of the carbonyl group, thereby reinforcing the strength of the hydrogen bond, while the strong electron-withdrawing inductive effect of the formyl group might result in the reverse way. In the case of benzamide and α-phenylacetamide, the E values in Table XV were obtained by dividing the difference of ΔH and $\Delta H'$ by two, because these molecules are connected to each other by two NH···O bonds per molecule in the crystalline state, as is revealed by Penfold and White[9].

The value of E for *p*-aminoacetophenone, 2.8 kcal./mol.(=5.6/2 kcal./mol.), is somewhat smaller than those of the compounds in the second category. This is probably due to the separation of the amino and carbonyl groups by a phenyl nucleus which might prevent the mutual approach of molecules to some extent, although the positive charge on the NH group as well as negative charge on the C=O group are presumably increased to favor the formation of a strong NH···O bond due to the existence of a phenyl nucleus. An alternative way of explaining the smaller value of the energy of hydrogen bond in this compound is that it may behave as an amine forming hydrogen bonds of NH···N type in the crystalline state. This problem must be solved either by infrared absorption measurement or by crystallographic analysis.

The energies of an NH···O bond estimated for δ-valerolactam and ε-caprolactam amount to 6~7 kcal./mol. and can be favorably compared with those observed for carboxylic acids[15] which are known to form dimers in the crystalline state as well as in liquid and in solutions. Here, it would be reasonable to assume that the two lactams exist as dimers in the crystalline state, because their molecular structures are such as to favor the formation of dimers (N–H is located in the *cis* position to C=O), just like those of carboxylic acids, although the crystal structures of the lactams are not analyzed as yet. Thus, the appreciable increase in the strength of NH···O bonds in

11) H. M. Randall et al., "Infrared Determination of Organic Structures", D. Van Nostrand Co., Inc., New York (1949), pp. 82, 147.
12) H. Baba and T. Suzuki, *J. Chem. Soc. Japan, Pure Chem. Sec. (Nippon Kagaku Zasshi)*, **81**, 366 (1960).
13) K. Kurosaki, ibid., **79**, 1339, 1362 (1958).
14) C. P. Smyth, "Dielectric Behavior and Structure", McGraw-Hill Book Co., Inc., New York (1955), p. 237.
15) F. H. McDougall, *J. Am. Chem. Soc.*, **58**, 2585 (1936); E. W. Johnson and L. K. Nash, ibid., **72**, 547 (1950); R. E. Lundin, F. E. Harris and L. K. Nash, ibid., **74**, 743, 4654, (1952).

these cases, compared with a "normal" NH···O bond, can be understood as due to the conjugation of π-electrons in the ring system formed by two molecules:



A simple treatment of this problem based on the assumption of a dipole-dipole interaction between hydrogen-bond-forming molecules is mentioned in a previous paper[7]. It should be mentioned that the estimated value of the energy of an NH···O bond for δ-valerolactam, 6.3 kcal./mol., is consistent with Tsuboi's value, 5.2 kcal./mol., obtained by the infrared spectroscopy on the association of the molecules in a CCl_4 solution[16].

In relation to the formation of ring dimers by the lactams mentioned above, a comment will be made on the molecular configuration of formanilide which is regarded as forming a "normal" hydrogen bond in this paper. Here it was implicitly assumed that the configurational position of the NH group was *trans* to the C=O group in this molecule, following the experimental result on the molecular structure of acetanilide[9]. Recently, however, Suzuki[17] has observed the existence of a *cis* form of this molecule in a CCl_4 solution in the study of its infrared absorptions and dipole moment: *cis* and *trans* forms exist in equilibrium, the energy difference being 600 cal./mol. The problem is whether the two forms should be found even in the crystalline state or not. Considering the "normal" value of the energy of hydrogen bond for formanilide, 4.1 kcal./mol., estimated in this study, the author is inclined to conclude that the molecule exists in the crystalline state as a *trans* form, because, if there is an appreciable amount of *cis* form, forming ring dimers, the energy of the hydrogen bond should be much higher, as can be seen from the values estimated for δ-valero- and ε-caprolactam.

Comparison of the Energies of Hydrogen Bonds in Crystalline State between the Types NH···O and OH···O.—In order to get a general view on the strength of NH···O and OH···O bondings in the crystalline state, a comparison was made between the two types of the hydrogen bonds using the energy values of the bonds estimated by the author in series of studies[5], together with those published already by other investigators (recalculated by the author). The result is shown in Fig. 1. In this figure, numbers of the compounds which show particular values of the energy of the hydrogen bond are plotted against the energy values: for the sake of simplicity the values are rounded off to integers approximating to the original values. The following compounds and energy values are picked up mostly from the publications by Seki and his collaborators[18]: CH_3OH 5.9 (6.0), C_2H_5OH 6.5 (5.8), n-C_3H_7OH 6.4 (5.6), n-C_4H_9OH 6.0 (5.1), n-$C_5H_{11}OH$ 6.5 (5.5), cyclohexanol 5.8 (3.5), pentaerythritol 5.4 (4.0), *i*-erythritol 5.4[19] (4.0), phenol 4.5 (5.2), hydroquinone 4.7 (4.2), ice 4.4 (4.7), hydrogen peroxide 4.5 (5.1), diformylhydrazine 6.0[1] (6.6), diacetylhydrazine 3.1[8] (4.8), diketopiperazine 5.3[6] (4.9), urea 3.5[4] (3.4), oxamide 4.2[3] (4.0), sulphamide 4.7[3] kcal./mol., respectively. The figures in parentheses are the energy values estimated by the procedure proposed by the author[2]. Acetamide 3.1, *N*-methylacetamide 3.4[2], acetylglycine-*N*-methylamide 3.2[20] and acetylglycineanilide 2.5[20] kcal./mol., respectively, are also included in Fig. 1.

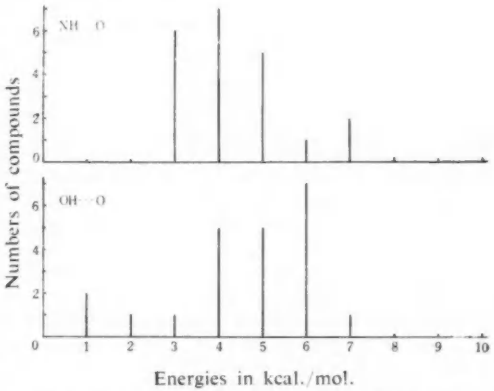


Fig. 1. Comparison of the energy of hydrogen bond between the type NH···O and OH···O.

It may be concluded from the distribution of the energy of hydrogen bonds as is seen in the figure, that the "normal" value of the energy of the NH···O bond can be 4 ± 1 kcal./mol., while that of the OH···O bond amounts to 5 ± 1 kcal./mol., and that a higher value both for NH···O and OH···O bonds can be ascribed to a favorable effect of conjugation in the π electron systems as is seen in the formation of ring dimers by lactams and carboxylic acids. The lower values observed in the case of

16) M. Tsuboi, This Bulletin, 24, 75 (1951).
17) I. Suzuki, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **80**, 353 (1959).
18) S. Seki et al., "Hydrogen Bond", Iwanami, Tokyo (1956), p. 41.
19) A. Shimada, This Bulletin, 32, 330 (1959).
20) A. Aihara, *J. Chem. Soc. Japan, Prue Chem. Sec.* (*Nippon Kagaku Zasshi*), **76**, 495 (1955).

OH···O bonding are due to the steric hindrance by substituted groups which prevent the mutual approach of hydroxyl groups.

Summary

The energies of hydrogen bonds formed in the crystalline state of thirteen amides and one amine have been estimated following the procedure proposed in a previous paper. The values of the energy of an NH···O bond can be estimated at about 4±1 kcal./mol. for the majority of compounds treated in this paper, except two lactams which show energy values as large as 6~7 kcal./mol. A comparison was made between the energy values of NH···O and OH···O bonding using the data obtained by the author and other investigators.

Department of Chemistry
Faculty of Liberal Arts and Science
Shinshu University
Agata-machi, Matsumoto

Effect of Compacting and Grinding on the Hydrogen Reduction of Nickel Oxide

By Yoshio IIDA and Kenzo SHIMADA

(Received February 10, 1960)

It has long been known that metal is hardened as a result of plastic deformation when it is pressed or rolled. Recently, it has been reported that the physical or chemical properties of non-metallic substances such as oxide[1], ionic crystal[2] and mineral[3] also vary when they are ground or pressed. Although these phenomena are not explained completely at present, increasing interest is being shown in the plastic deformation of ionic crystal.

In the previous paper[4] dealing with the hydrogen reduction of nickel oxide, it was shown that the rate of reduction was affected by the crystal distortion induced by the dissolution of minor additives. It is thought to be desirable to see how the reduction of nickel oxide is affected by grinding or compacting. In the present paper the isothermal reduction of nickel oxide, in a single crystal form or in powder form, subjected to grinding or compacting has been studied, consideration being given to the surface area of the specimens.

Experimental

Chemical pure nickel sulfate recrystallized four times was heated at 850°C for 10 hr. to obtain nickel oxide powder. The average unit-particle size of nickel oxide was about 0.5 μ according to its electron micrograph. Nickel oxide was compacted into tablets at various pressures of 0 to 2000 kg./cm^2. A single crystal of nickel oxide was prepared by the method similar to that described by Verneuil[5].

The rate of reduction by hydrogen was determined by a thermobalance with a sensitivity of 0.5 mg. and could be used in hydrogen and in vacuo. Since the rate of reduction of nickel oxide by hydrogen at 200°C was a suitable magnitude for a comparison of the rates of specimens subjected to various treatments, the isothermal reduction was carried out at 200°C in the present study. In all cases of this experiment, nickel oxide of a constant weight of 500 mg. was heated up to 200°C in a vacuum of 10^{-3} mmHg and reduced at a constant temperature of 200°C by hydrogen whose flowing velocity was 100 cc./min.

Hydrogen from a commercially available cylinder was used after being purified with concentrated sulfuric acid, hot platinum asbestos, potassium hydroxide and phosphorus pentoxide.

The surface area of nickel oxide powder was determined by the B. E. T. method[6] in which nitrogen was the absorbent.

Results and Discussion

Effect of Compacting Pressure.—The bulk density of nickel oxide pressed at various pressures of 0 to 2000 kg./cm^2 was measured from the ratio of weight of dimension. Table I shows that the bulk density increases and the calculated porosity decreases with an increase

1) J. B. Wachtman, *J. Am. Ceram. Soc.*, **37**, 291 (1954). A. S. Keh et al., *Acta Metallurgica*, **7**, 694 (1959).
2) S. Sonoike and T. Ninomiya, *Metal Physics* (*Kinzoku-Buturi*), **1**, 67 (1955); H. Kanzaki, ibid., **1**, 108 (1955).
3) H. Takahashi, This Bulletin, **32**, 235, 245, 252 (1958).
4) Y. Iida and K. Shimada, ibid., **33**, 8 (1960).
5) J. Nakazumi, *J. Chem. Soc. Japan. Ind. Chem. Sec.* (*Kogyo Kagaku Zasshi*), **59**, 1304 (1956).
6) S. Brunauer, P. H. Emmett and E. Teller, *J. Am. Chem. Soc.*, **60**, 309 (1938).

TABLE I. BULK DENSITY, POROCITY AND SURFACE AREA OF NICKEL OXIDE PRESSED AT VARIOUS PRESSURES

Compacting pressure, kg./cm²	0	200	500	1000	2000
Bulk density, g./cc.	1.7	2.1	2.5	3.0	3.4
Surface area, m²/g.	1.60	—	1.54	1.67	1.50
Porosity, %	75	69	63	56	50

in the compacting pressure. Although the surface area could be thought to decrease with an increase in the pressure, it does not vary with the pressure as shown in Table I. From these data it is found that a contact point which is formed by an atomic distance between nickel oxide particles and which prevents the adsorption of nitrogen, does not decrease with the pressure in spite of the decreasing porosity. In the previous paper[7] dealing with the oxidation of nickel powders it was shown that a surface area of nickel powders with a ductile property did not vary with the compacting pressure in the range of 0 to 1000 kg./cm². In view of the results concerning nickel powders, the results shown in Table I can be accepted as correct.

The effect of compacting pressure on the isothermal reduction of nickel oxide is shown in Fig. 1. It is seen that the rate of reduction increases with an increase in the compacting pressure below 1000 kg./cm². It should be noted that the amount of reduced nickel oxide in 5 hr. at 200°C is only 50% for an unpressed nickel oxide whereas it is 100% for nickel oxide pressed at 1000 kg./cm². Inasmuch as the surface area does not vary with the compacting pressure as shown in Table I, the accelerated reduction by compacting can be attributed to the crystal distortion which may be induced by compacting. Although the mechanism of hydrogen reduction of nickel oxide compact is very complicated[8] and not yet understood clearly, the porosity of the compact is thought to play an important role in the reduction which is a chemical reaction of solid with gas. The rate is lower at 2000 kg./cm² than at 1000 kg./cm². It is found that the effect of porosity on the reduction exceeds that of compacting when the pressure is relatively high and the porosity is small.

From these data it is expected that the rate becomes higher when nickel oxide compact pressed at a high pressure is ground into fine powders and its porosity becomes larger. Fig. 2 shows that the reduction is promoted further when the compact pressed at 1000 kg./cm² is ground into powders. It is seen that the rate of reduction of nickel oxide powder prepared by compacting and grinding is obviously higher than that of the nickel oxide powder subjected to no mechanical treatment and having the same porosity. This means that the effect of mechanical treatment on the hydrogen reduction is large and the effect of the porosity is small.

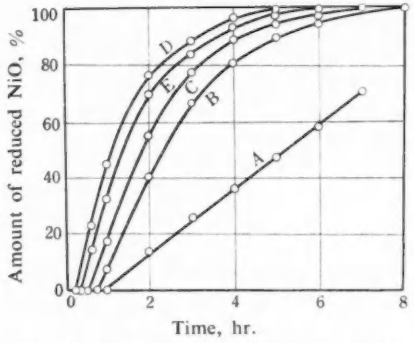


Fig. 1. Isothermal reduction of nickel oxide pressed at various pressures. The reduction temperature is 200°C. A is for nickel oxide unpressed, B, for nickel oxide pressed at 200 kg./cm², C, at 500 kg./cm², D, at 1000 kg./cm², and E, at 2000 kg./cm², respectively.

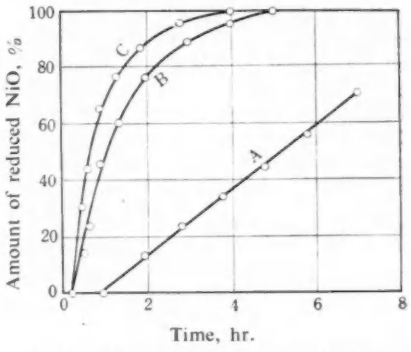


Fig. 2. Effect of compacting and grinding on the hydrogen reduction of nickel oxide at 200°C. A is for nickel oxide without compacting, B, nickel oxide compacted at 1000 kg./cm², and C, nickel oxide ground after compacting at 1000 kg/cm².

7) Y. Iida, S. Ozaki and K. Shimada, *Reports of the Government Industrial Research Institute, Nagoya (Nagoya Kogyo Gijitsu Shikenjo Hokoku)*, **10**, 569 (1957).

8) Y. Iida and K. Shimada, This Bulletin, to be published.

Effect of Grinding.—Isothermal reductions of a single crystal of nickel oxide and of nickel oxide powders prepared by grinding of the crystal were compared with that of nickel oxide prepared by ignition of nickel sulfate. The results in Fig. 3 show that the rate of

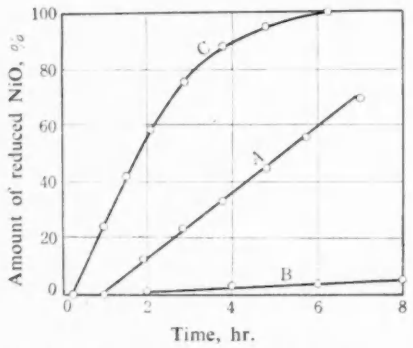


Fig. 3. Effect of grinding of a single crystal on the hydrogen reduction of nickel oxide at 200°C. A is for nickel oxide powders from ignition of nickel sulfate, B, a single crystal of nickel oxide and C, nickel oxide powders from grinding of the single crystal.

reduction of nickel oxide prepared by grinding of the single crystal is larger than those of a single crystal and nickel oxide prepared by ignition of nickel sulfate. It is a natural result from the consideration of the surface area that the rate of reduction of nickel oxide powders prepared by grinding of the single crystal exceeds that of the single crystal. The particles size of nickel oxide prepared by grinding of the single crystal is about 50 μ by a microphotograph and its surface area is too small to be determined by the B. E. T. method. The particle size of nickel oxide prepared by ignition of nickel sulfate is about 0.5 μ and is evidently smaller than that of nickel oxide prepared by grinding of the single crystal. The rate of reduction of nickel oxide prepared by ignition, however, is lower than that of nickel oxide prepared by grinding. As the single crystal was prepared from nickel oxide by igniting C. P. nickel sulfate, the purity of both nickel oxides is thought to be essentially the same. From the viewpoint of the surface area of nickel oxides, these facts cannot be understood reasonably.

Recently, it was reported that a change in physical properties of sodium chloride[2], zinc oxide* and kaoline[3] by grinding or compressing might be attributed to the crystal distortion which was caused by the plastic deformation. From these considerations it is concluded that the accelerated reduction of nickel oxide prepared by grinding of the single crystal results from the crystal distortion induced by grinding. On the basis of these data, it seems reasonable to assume that nickel oxide, either in powder form or single crystal form, shows plastic deformation which results in the accelerated hydrogen reduction, when it is compacted or ground.

Summary

The rate of hydrogen reduction of nickel oxide increases with an increase in the compacting pressure although the bulk density of nickel oxide increases and the surface area does not vary with the pressure. The rate of reduction of nickel oxide which is prepared by the grinding of a single crystal and has a large particle size of 50 μ is higher than that of nickel oxide which is prepared by ignition of nickel sulfate and has small particle size of 0.5 μ. It seems reasonable to conclude that the accelerated reduction is attributed to the crystal distortion which may result from plastic deformation by grinding or compacting.

The authors wish to express their sincere thanks to Tochigi Kagaku Kogyo Co. for the presentation of the single crystal.

Government Industrial Research Institute
Kitaku, Nagoya

* It is easily observed that zinc oxide turns yellow when it is ground after sintering.

A Molecular Orbital Theory of Saturated Compounds. I. Ionization Potential and Bond Dissociation Energy

By Kenichi FUKUI, Hiroshi KATO and Teijiro YONEZAWA

(Received February 10, 1960)

Recently, several simple quantum-mechanical methods have been proposed to discuss physical properties of saturated compounds. That is to say, the bond energy and the heat of formation have been studied by Brown's bond orbital method[1] and by Dewar's perturbation treatment[2]. The ionization potential and the electron affinity have been investigated by the equivalent orbital method of Lennard-Jones and Hall, Franklin's united atom treatment[4], and Higuchi's UAO method[5].

In the present paper, the aliphatic and alicyclic compounds will essentially be treated by the use of the method previously presented by Yoshizumi[6], in which the skeleton consisting of carbon chain in an aliphatic molecule was exclusively taken into consideration. Thus, the ionization potential, heat of formation and bond energy of various saturated compounds are calculated according to this procedure, and the results obtained are in good agreement with experiments.

Theoretical

Yoshizumi applied the strategy which was usual for treating the π-electron system to attacking the σ electron system. Namely, in treating aliphatic molecules the sp^3-hybridized orbital of a carbon atom appeared in place of the $2p\pi$ atomic orbital in the treatment of conjugated molecules. He neglected the C–H bond of paraffin hydrocarbons for the sake of simplicity. Thus the molecular orbital was represented by a linear combination of sp^3 hydridized orbitals of carbon atoms, as follows:

$$\phi_j = \sum_j C_{rj}\phi_r \qquad (1)$$

where C_{rj} is the coefficient of the rth atomic sp^3 hybridized orbital of carbon, ϕ_r, in the jth molecular orbital, ϕ_j. All the overlap integrals were neglected as in the simple LCAO MO theory in π-electron system. The resonance integrals in C–C bond and that between the orbitals attached to the same carbon atom were assumed as

$$\int \phi_i H \phi_{i+1} d\tau = \beta$$

and

$$\int \phi_i H \phi_{i+1} d\tau = m_i\beta$$

respectively, where m was a small quantity which was determined by Yoshizumi as $m^2 = 0.12$ referring to the value of dipole moment of halogen-substituted hydrocarbons. The resonance integrals between non-neighboring atomic orbitals were neglected. In Fig. 1, it will be seen that when the number of an orbital i is odd, the integral corresponds to that between two adjacent carbon atoms, and when i is even, the resonance integral between two atomic orbitals of one and the same carbon atom.

In Yoshizumi's paper only linear paraffins were treated. In treating the other saturated compounds, it is necessary to introduce some emprical parameters besides.

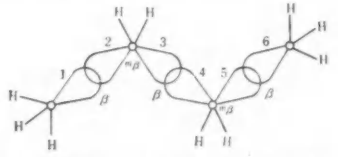


Fig. 1. Carbon sp^3 hybridized orbitals in *n*-butane.

Parameters

For tertiary and quaternary carbon atoms in branched chain hydrocarbons, the Coulomb integral is tentatively taken as $\alpha_C = \alpha - \delta\beta$ in common according to their electronegativity, where α is the Coulomb integral of sp^3 hybridized orbital of primary and secondary carbon atoms. The value of δ is determined as equal to 0.3 considering the value of ionization potential of *i*-butane.

For alkyl halides (R–X), the Coulomb integral, α_X, is taken as $\alpha + a_i\beta$ and the resonance integral, β_{C-X} as equal to $b_i\beta$, where b is determined by the formula $\frac{\beta_{C-X}}{\beta_{C-C}} = \frac{S_{C-X}}{S_{C-C}}$ (S is

1) R. D. Brown, *J. Chem. Soc.*, **1953**, 2615.
2) M. J. S. Dewar and R. Pettit, *J. Chem. Soc.*, **1954**, 1625.
3) J. Lennard-Jones and G. G. Hall, *Disc. Faraday Soc.*, **10**, 18 (1951); G. G. Hall, *Proc. Soc.*, **A213**, 102 (1952), *Trans. Faraday Soc.*, **49**, 113 (1953).
4) J. L. Franklin, *J. Chem. Phys.*, **22**, 1304 (1954).
5) J. Higuchi, *Summary of Electronic Structure Symposium*, 23 (1958).
6) H. Yoshizumi, *Trans. Faraday Soc.*, **53**, 125 (1957).

obtainable from Mulliken's table[7]*) and a is determined referring to the values of ionization potential of C–X bonding electron of methyl halides assigned by McDowell[8]. The values of a and b used in the present paper are listed in the following table.

X	a	b
F	0.91	0.51
Cl	0.56	0.65
Br	0.54	0.58
I	0.50	0.53

For alkyl alcohols, the Coulomb integral of alcoholic oxygen, α_{OH}, is taken as $\alpha+0.5\beta$ and the resonance integral, β_{C-OH}, as 0.5β, where β_{C-OH} is determined by using the formula stated above (the value of S_{C-OH} is put nearly equal to that of S_{C-O}). Since the value of the ionization potential of C–OH bonding electron of methyl alcohol has never been known, the value of α_{OH} is tentatively assumed as above.

Results and Discussion

a) Ionization Potential.—(1) It was pointed out[5,9] that in normal parraffins their first ionization could be attributed to the removal of one electron localized in C–C bonds. This was supported by the difference between observed C–H and C–C bond energies, and also by the result of Hall's calculation[3]. Namely, he elucidated the theory that in the highest occupied orbital of normal paraffins the electron density of C–C bond was larger than that of C–H bond. Therefore, the ionization potential may safely be considered to be equal to the negative of the highest occupied orbital energy obtained by solving the secular determinant. The calculated results are listed in Table I, and the plot of calculated energy values versus observed ionization potentials is shown in Fig. 2. From the linear relation between the calculated and the observed, the values of α and β are estimated as -5.850 eV. and -6.364 eV., respectively. With these values, the recalculated ionization potentials of paraffins are listed in the third column of Table I which are in good agreement with the observed, except in ethane.

TABLE I. CALCULATED ENERGIES OF THE HIGHEST OCCUPIED ORBITAL AND OBSERVED IONIZATION POTENTIALS OF NORMAL PARAFFINS

Compound	Calcd. energy[a] λ_{ho}	ε_{ho} eV.	Obsd. I. P.[9] eV.
Ethane	1.0000	12.21	11.76
Propane	0.8444	11.22	11.21
n-Butane	0.7752	10.77	10.80
n-Pentane	0.7388	10.55	10.55
n-Hexane	0.7170	10.41	10.43
n-Heptane	0.7034	10.33	10.35
n-Octane	0.6921	10.25	10.24
n-Decane	0.6820	10.19	10.19

a) λ_{ho} is the coefficient in $\varepsilon_{ho}=\alpha+\lambda_{ho}\beta$, where ε_{ho} is the highest occupied orbital energy.

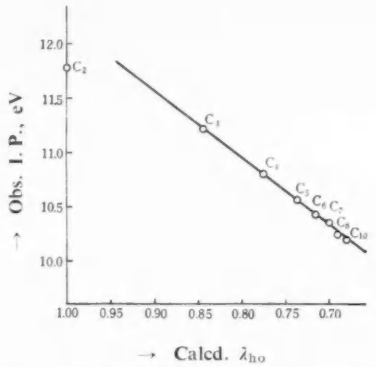


Fig. 2. Ionization potential of normal paraffins vs. λ_{ho}.

(2) In the calculation of cycloparaffins the steric configuration is left out of consideration, all the C–C–C angles being assumed equal to that of linear molecules. Therefore, no parameter is introduced anew. The results are listed in Table II, together with observed and Franklin's values[4] for the sake of comparison. It is seen that the smaller the deviation of an actual C–C–C angle from the normal one, the better agreement between the observed and the calculated results is found. This suggests that there is some room for improvement in future.

(3) For branched chain paraffins, the calculated values of ionization potentials are listed in the third column of Table III.

(4) It is considered that the first ionization potential of alkyl alcohols is due to removal of one of the lone pair electrons of the oxygen atom. The correlation of calculated values of the highest occupied orbital energies with observed ionization potentials is examined and an apparent linear relation between them is shown in Table IV, and it may result from the following cause. In general, the inductive effect of alkyl group may exert an influence

* The evaluation of the overlap integral, S_{C-X}, for bromine using the Slater orbital can not be carried out in the usual way. Therefore, the nearest integral power of r in Slater orbital is adopted in our calculation.

7) R. S. Mulliken, C. A. Rieke, D. Orloff and H. Orloff, *J. Chem. Phys.*, **17**, 1248 (1949).

8) D. C. Frost and C. A. McDowell, *Proc. Roy. Soc.*, **A241**, 194 (1957).

9) R. E. Hornig, *J. Chem. Phys.*, **16**, 105 (1948).

TABLE II. CALCULATED ENERGIES OF THE HIGHEST OCCUPIED ORBITAL AND OBSERVED IONIZATION POTENTIALS OF CYCLOPARAFFINS

Compounds	Calcd. energy λ_{ho}	Calcd. energy ε_{ho} eV.	Calcd. I. P.[a] by Franklin eV.	Obsd. I. P. eV.	Deviation[b] from normal C-C-C angle
Cyclopropane	0.8806	11.45	11.76	10.23[a]	+24°44′
Methylcyclopropane	0.7050	10.34	10.98	9.88[a]	—
Cyclobutane	0.6600	10.05	10.21		+ 9°44′
Cyclopentane	0.7520	10.64	10.80	11.10[a]	+ 0°44′
Cyclohexane	0.6600	10.05	10.21	10.40[a]	− 5°16′

a) J. Hissel, *Bull. Soc. Roy. Sci. Liege*, **21**, 4576 (1952).
b) L. F. Fieser and M. Fieser, "Textbook of Organic Chemistry", D. C. Heath Co., Boston (1950).

TABLE III. CALCULATED ENERGIES OF THE HIGHEST OCCUPIED ORBITAL AND OBSERVED IONIZATION POTENTIALS OF BRANCHED PARAFFINS

Compounds	Calcd. energy λ_{ho}	Calcd. energy ε_{ho} eV.	Obsd. I. P. eV.
Methylpropane	(0.6995)	(10.30)	10.40[a]
Tetramethylmethane	0.6995	10.30	10.30[b]
Ethylpropane	0.6469	9.96	10.19[c]
2,2-Dimethylbutane	0.6378	9.93	9.90[c]
3-Methylpentane	0.6222	9.81	9.81[c]

a) D. P. Stevenson, *J. Am. Chem. Soc.*, **64**, 2766 (1942).
b) D. P. Stevenson, *Disc. Faraday Soc.*, **10**, 35 (1951).
c) M. B. Koffel, *J. Chem. Phys.*, **17**, 44 (1948).

TABLE IV. CALCULATED ENERGIES OF THE HIGHEST OCCUPIED ORBITAL AND TOTAL ELECTRON DENSITIES AT OXYGEN AND OBSERVED IONIZATION POTENTIALS OF ALKYL ALCOHOLS

Compounds	Calcd. energy λ_{ho}	Calcd. total dens. at oxy.	Obsd. I. P. eV.
Methyl alcohol	(0.8090)	1.4472	10.95[a]
Ethyl alcohol	0.7589	1.4618	10.60[a]
Propyl alcohol	0.7350	1.4643	10.46[a]
Butyl alcohol	0.7180	1.4712	10.30[b]
i-Propyl alcohol	0.6807	1.4784	10.27[b]

a) J. D. Morrison and A. J. C. Nicholson, *J. Chem. Phys.*, **20**, 1021 (1952).
b) Reference 10).

on the electronegativity of oxygen lone pairs[10]. Therefore, the parallelism may be more directly understood by comparing the observed ionization potential with the magnitude of total σ electron density the oxygen atom which is also given in the same table.

10) I. Omura, H. Baba and K. Higashi, This Bulletin, **28**, 147 (1955).

b) Total Electronic Energy and Bond Dissociation Energy.—The calculated total electronic energies are compared with observed values of heat of formation of several normal paraffins. The results indicate a good parallelism between them. For each increase of one CH_2-group the increment in the calculated energy lies between $2\alpha+2.057\beta$ and $2\alpha+2.059\beta$, while the observed difference in energy is in the range $-4.6\sim-5.2$ kcal./mol.[1] as shown in Table V. For cycloparaffins, the calculated total electronic energies are compared with observed heats of combustion and both the values per CH_2 are listed in Table VI*. Thus we can conclude that a simple additivity rule may approximately hold in normal paraffins and cycloparaffins.

In order to calculate the bond dissociation energy, we assume that it can be obtained by the following formula**.

$$D(R_1-R_2)=\{E(R_1\cdot)+E(R_2\cdot)\}-E(R_1-R_2) \quad (2)$$

where $D(R_1-R_2)$ denotes the dissociation energy of the bond between R_1 and R_2 in the compound R_1-R_2, E is the total electronic energy, and R_1 and R_2 are two alkyl groups. For simplicity, the energy of an odd electron in alkyl radical, which is localized on one carbon atom, is assumed equal to α. The results are listed in Table VII, where the numbers in parentheses are the observed values of bond energy and the numbers below them are the calculated values of bond energy in units of $(-\beta)$. In lower paraffins, the values obtained by Eq. (2) are shown to be a good measure for bond energy value. In hexane, a slight disagreement is observed. No internuclear repulsion is involved in the present treatment, and hence this discrepancy may, at

* The failure seen in the case of cyclobutane may probably be attributed to the approximation adopted in which steric circumstances are entirely disregarded.
** In the calculation of energy of alkyl radicals, the effect of hyperconjugation of alkyl group is left out of account.

TABLE V. CALCULATED TOTAL ELECTRONIC ENERGIES AND OBSERVED HEATS OF FORMATION OF NORMAL PARAFFINS

Compounds	Calcd. total energy	Obsd. heat of formation[a] kcal./mol.	Difference Calcd.	Obsd. kcal./mol.
Ethane	$4\alpha+2.000\beta$	−20.2		
			$2\alpha+2.057\beta$	−4.6
Propane	$6\alpha+4.057\beta$	−24.8		
			$2\alpha+2.058\beta$	−5.0
n-Butane	$8\alpha+6.115\beta$	−29.8		
			$2\alpha+2.058\beta$	−5.2
n-Pentane	$10\alpha+8.157\beta$	−35.0		
			$2\alpha+2.058\beta$	−5.0
n-Hexane	$12\alpha+10.231\beta$	−40.0		
			$2\alpha+2.059\beta$	—
n-Heptane	$14\alpha+12.290\beta$	—		

a) F. D. Rossini et al., "Selected Values of Physical and Thermodynamic Properties of Hydrocarbons and Related Compounds", Carnegie Press, Pittsburgh (1953).

TABLE VI. CALCULATED TOTAL ELECTRONIC ENERGIES AND OBSERVED HEATS OF COMBUSTION OF CYCLOPARAFFINS

Compounds	Calcd. energy Total	Calcd. energy per CH_2	Obsd. heat of combustion[a] Total kcal./mol.	Obsd. heat of combustion[a] per CH_2 kcal./mol.
Cyclopropane	$6\alpha+6.203\beta$	$2\alpha+2.067\beta$	505.5	168.5
Cyclobutane	$8\alpha+8.225\beta$	$2\alpha+2.056\beta$	662.5	165.5
Cyclopentane	$10\alpha+10.293\beta$	$2\alpha+2.059\beta$	797	159
Cyclohexane	$12\alpha+12.348\beta$	$2\alpha+2.058\beta$	939	157

a) Reference b) in Table II.

TABLE VII. DISSOCIATION ENERGIES OF C–C BONDS

R_1 \ R_2	Methyl	Ethyl	Propyl	Butyl
Methyl	(82) 2.000	(82) 2.057	(79) 2.058	(78) 2.058
Ethyl	—	(82) 2.115	(79) 2.116	(78) 2.116
Propyl	—	—	(76) 2.117	—

Ethane $C\overset{(82)}{\underset{2.000}{}}C$

Propane $C\overset{(82)}{\underset{2.057}{}}C-C$

n-Butane $C\overset{(78)}{\underset{2.058}{}}C\overset{(82)}{\underset{2.115}{}}C-C$

n-Pentane $C\overset{(78)}{\underset{2.058}{}}C\overset{(79)}{\underset{2.116}{}}C-C-C$

n-Hexane $C\overset{?}{\underset{2.058}{}}C\overset{(78)}{\underset{2.116}{}}C\overset{(76)}{\underset{2.117}{}}C-C-C$

TABLE VIII. CALCULATED AND OBSERVED BOND DISSOCIATION ENERGIES OF ETHYL HALIDES

X in C_2H_5X	C–X bond diss. energy Calcd. $(-\beta)$	Obsd.[a] (kcal./mol.)
F	2.690	107.0
Cl	2.049	66.5
Br	1.896	54·0
I	1.749	45.5

a) Reference b) in Table II.

least partly, be ascribed to such a rough approximation adopted. Further improvement should be made on this point.

With regard to ethyl halides, the bond dissociation energies of C–X bond are obtained and compared with experimental values in Table VIII. Compounds are arranged in decreasing order of bond energies. It is seen that the result coincides well with experiments.

Summary

The ionization potential and the bond dissociation energy of saturated compounds are calculated by Yoshizumi's C–C bond skeleton method. With regard to ionization potentials, the calculated results are parallel to the observed values. In particular, for normal paraffins the parallelism is almost complete. For total electronic energies, the calculated results indicate a simple additivity rule in normal and cycloparaffins, and for bond dissociation energies an approximate agreement is obtained between the calculated and the observed values.

In spite of the rough approximation involved in the present treatment, it may be said that the results obtained are rather satisfactory.

Faculty of Engineering
Kyoto University
Sakyo-ku, Kyoto

A Molecular Orbital Theory of Saturated Compounds. II. Chemical Reactivity in Nucleophilic Substitution

By Kenichi FUKUI, Hiroshi KATO and Teijiro YONEZAWA

(Received February 26, 1960)

It is well known that a nucleophilic substitution of aliphatic compounds has available two reaction mechanisms, one being the unimolecular (S_N1) and the other the bimolecular mechanism (S_N2)[1]. The explanation of reactivity in aliphatic substitution has hitherto been carried out only by means of empirical electronic theory. No quantum-mechanical treatment of chemical reactivity in nucleophilic aliphatic substitution has ever been put into practice though some theoretical studies have been made with regard to the Walden inversion by Jaffé[2] and others[3,4].

In Part I[5] of this series of papers the ionization potential and the bond dissociation energy of several aliphatic compounds were calculated by the use of the C-C skeleton method proposed by Yoshizumi[6], resulting in a good agreement with the experiments.

In the present paper, an attempt is made to elucidate the reactivity in nucleophilic substitution of some alkyl halides, following the same quantum-mechanical procedure as used in Part I[5]. Also by the use of the procedure of simple LCAO MO method of conjugated molecules the stability of some aliphatic ions is discussed, in which the hyperconjugative effect is thought to be prevailing.

S_N2 Reaction

The reaction of this type is expressed by the equation[1]

$$A^- + R-X \rightarrow [A\cdots R\cdots X]^- \rightarrow A-R + X^-$$

where A^- is a nucleophilic reagent, R is an alkyl group, X is the group to be expelled, and the compound in brackets, stands for the activated complex.

The structure of activated complex is schematically written in Fig. 1. In this state, R′, R″, R‴ and C-atom lie all in one plane, and also A, C-atom and X lie on one straight line perpendicular to the plane. Four sp^3 orbitals of the central carbon atom may change to three sp^2 orbitals and one p orbital which binds both A and X to C-atom. However, for the sake of simplicity, we neglect the influence of altered hybridization at the central C-atom and adopt the same values for Coulomb and exchange integrals as sp^3 hybridization. The overlap integral between sp^2 orbital and p orbital is put equal to zero. The contribution of the reagent A to the energy of the total system is regarded as the same in all the reactions of the same type. We assume that in the transition state two σ electrons in the C–X bond to be broken are localized at X-atom*. According to the present model, the difference in the energy between the transition state and the initial state, ΔE, is calculated as

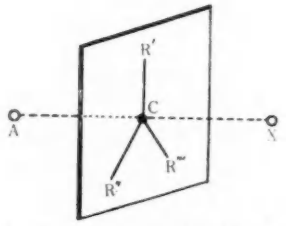


Fig. 1. Transition state in S_N2 reaction.

$$\Delta E = \{E[R^+] + E[X^-]\} - E[R-X] \qquad (1)$$

where $E[A]$ is the calculated total energy of the molecule or ion A. This procedure corresponds to the localization method in the theory of aromatic substitution. This energy, ΔE, is considered to be a measure of the activation energy of the reaction.

Calculations are made with respect to some alkyl bromides, chlorides and iodides. The second order reactions of ethoxide anion and

1) C. K. Ingold, "Structure and Mechanism in Organic Chemistry", Cornell University Press, Ithaca, N. Y. (1953), Chapter VII.
2) H. H. Jaffé, *J. Chem. Phys.*, **22**, 1618 (1953).
3) I. Dostrovsky, E. D. Hughes and C. K. Ingold, *J. Chem. Soc.*, **1946**, 173.
4) S. Glasstone K. J. Laidlar and H. Eyring, "The Theory of Rate Processes", McGraw-Hill Book Co., New York, London (1941), pp. 135, 400.
5) U. Fukui, H. Kato and T. Yonezawa, This Bulletin, **33**, 1197 (1960).
6) H. Yoshizumi, *Trans. Faraday Soc.*, **53**, 123 (1957).

* Two alternative models may be possible. The one is the model in which the two σ electrons are localized in the C-X bond, and the other is that in which the two electrons move in A···C and C···X bonds. Whichever of these two models one may adopt, only a constant term will be added to the energy difference obtained in the present treatment and no serious alteration arises with regard to the results.

chloride anion with alkyl bromides are investigated and the calculated results are compared with the ovserved activation energy and reaction rate in Table I. The agreement with this experiment is found satisfactory. In the present calculation, the energetic contribution of the attacking anion is not taken into account, since the anion is regarded as having the same energy in all the reactions of the same type.

In Table I, ΔE values of several alkyl bromides are listed for two types of nucleophilic reaction. In Table II, the experimental activation energies of alkyl iodides in the Menschutkin type reaction and in the substitution with the chloride anion are compared with the calculated values of ΔE. Table III indicates the results of calculation with regard to the reaction of alkyl chloride with iodide anion. It is observed that the value of ΔE increases according as the reaction rate decreases.

TABLE I

(1) $RBr + C_2H_5O^- \rightarrow C_2H_5OR + Br^-$ [1]

RBr	ΔE calc. $(-\beta)$	Activation energy obs. (kcal./mol.)	Rate constant obs. 10^3k at 55°C ($sec^{-1}mol^{-1}l.$)
C-Br	0.7420	20.0	34.4
C-C-Br	0.8156	21.0	1.95
C-C-C-Br	0.8174	—	0.55
C(C)-C-C-Br	0.8215	22.8	0.05
C(C)-C(C)-C-Br	0.9215	26.2	0.000008

(2) $RBr + Cl^- \rightarrow RCl + Br^-$ [1]

RBr	ΔE calc. $(-\beta)$	Activation energy obs. (kcal./mol.)	Relative rate obs.
C-Br	0.7420	15.8	100
C-C-Br	0.8156	17.9	1.97
C-C-C-Br	0.8174	18.1	1.29
C-C-C-C-Br	0.8174	—	1.29
C(C)-C-Br	0.8911	18.7	0.040
C(C)-C(C)-C-Br	0.9580	21.2	0.009

The calculated results in these tables indicate that the rate of S_N2 reaction decreases in the order methyl>ethyl>higher *n*-alkyl>*i*-propyl >*t*-butyl, in accordance with experience.

The absolute rate of S_N2 substitution is strongly affected by an exchange of the group to be displaced. The observed rate of hydrolysis or alcoholysis decreases in the order R–I >R–Br>R–Cl. Calculated values in brackets exhibit a similar tendency as follows:

C–C–I (0.7494$(-\beta)$)>C–C–Br(0.8156$(-\beta)$)
>C–C–Cl(0.9288$(-\beta)$).
C–C–C–I (0.7524$(-\beta)$)
>C–C–C–Br(0.8174$(-\beta)$)
>C–C–C–Cl(0.9305$(-\beta)$)

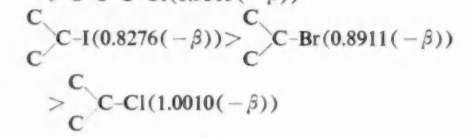


TABLE II

(3) $RI + Cl^- \rightarrow RCl + I^-$ [1]

RI	ΔE calc. $(-\beta)$	Activation energy obs. (kcal./mol.)	Relative rate Obs.
C-I	0.6880	16.0	100
C-C-I	0.7494	17.3	8.9
C-C-C-I	0.7524	17.3	5.2
C-C-C-C-I	0.7520	—	—
C(C)-C-I	0.8276	18.0	0.26

(4) $RI + (CH_3)_3N \rightarrow (CH_3)_3N^+R + I^-$ [1]

RI	ΔE calc. $(-\beta)$	Activation energy obs. (kcal./mol.)	Relative rate obs.
C-I	0.6880	9.7	100
C-C-I	0.7494	11.4	0.58
C(C)-C-I	0.8276	17.1	0.003

$RI + Et_3N \rightarrow RN^+Et_3 + I^-$ [a]

RI	ΔE calc. $(-\beta)$	Activation energy obs. (kcal./mol.)	Relative rate obs.
C-I	0.6880	9.7	100
C-C-I	0.7494	12.5	8.8
C-C-C-I	0.7524	—	1.7
C-C-C-C-I	0.7520	16.0	1.2
C(C)-C-I	0.8276	—	0.18

a) M. S. Newman, "Steric Effects in Organic Chemistry", John Wiley & Sons, Inc., New York (1956), p. 75.

TABLE III

$RCl + I^- \rightarrow RI + Cl^-$ [1]

RCl	ΔE calc. $(-\beta)$	Activation energy obs. (kcal./mol.)	Relative rate obs.
C-Cl	0.8480	—	93
C-C-Cl	0.9288	—	1.0
C-C-C-Cl	0.9305	—	0.37
C-C-C-C-Cl	0.9324	—	0.57
C(C)-C-Cl	1.0010	—	0.0076
C(C)-C(C)-C-Cl	1.0942	—	—

S_N1 Reaction

The rate determining step in this type of reaction is the ionic dissociation represented by the formula $RX \rightarrow R^+ + X^-$. Dewar[7] proposed a model for this transition state which may be written as S···R···X···S, where S stands for a solvent molecule, R is the alkyl group, and X is the group to be expelled. The factors affecting the rate of S_N1 reactions are considered to be (1) the rate of ionic dissociation of the molecule R–X, (2) the solvent effect, and so on. Refering to this we adopt as reactivity indexes the electron densities at the central C-atom (q_C) and at the X-atom (q_X) in the C–X bond. The larger q_C is, the faster is the rate of ionic dissociation of the molecule R–X. Furthermore, the larger q_X is, the stronger is the X···S binding.

Table IV

$RBr + OH^- \rightarrow ROH + Br^-$ [1]

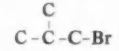
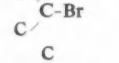
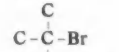
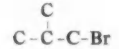
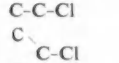
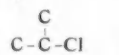
RBr	Calc. total density q_C	Calc. total density q_{Br}	Relative rate obs.
C-Br	0.5768	1.4231	1.00
C-C-Br	0.5979	1.4359	1.71
C-C-C-Br	0.6000	1.4401	—
	0.6005	1.4437	—
	0.6063	1.4511	44.7
	0.6056	1.4667	10^8
	0.6091	1.4874	—

Table V

(1) $R-X\ aq \rightarrow R^+\ aq + X^-\ aq$[7]

R-Cl	Calc. total density q_C	Calc. total density q_{Cl}	Heat of reaction by Evans, Q (kcal./mol.)
C-Cl	0.6076	1.3924	89
C-C-Cl	0.6219	1.4085	59
	0.6289	1.4223	37
	0.6257	1.4358	26

(2)

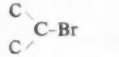
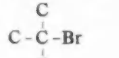
R-Br	Calc. total density q_C	Calc. total density q_{Br}	Heat of reaction by Evans, Q (kcal./mol.)
C-Br	0.5768	1.4231	86.5
C-C-Br	0.5979	1.4359	56.5
	0.6063	1.4511	34.5
	0.6056	1.4667	23.5

Table VI

Ions	Resonance energy calc. ($-\beta_0$)	Stabilization energy[10] obs. (kcal./mol.)
CH_3^+	0	0
$C_2H_5^+$	0.3350	36
$n\text{-}C_3H_7^+$	0.5081	—
$i\text{-}C_3H_7^+$	0.6462	66
$t\text{-}C_4H_9^+$	0.9043	84

Calculated values of q_C and q_{Br} are listed together with the rate of hydrolysis of alkyl bromide in Table IV in which a parallelism between them is found. Since there exist, however, no sufficient data available for comparison, the heat of reaction, Q, i. e. the approximate activation energy in S_N1 solvolysis of alkyl halides, computed by Evans[8,9] is used in Table V for this purpose. The magnitudes of q_C and q_X are seen to be parallel with the value of Q. Therefore, we conclude that q_C and q_X are useful reactivity indexes in S_N1 reaction.

Furthermore, it has been shown that the rates of some S_N1 reactions can be connected with stabilites of alkyl ions formed by dissociation[10,12]. These stabilities can be represented by the resonance energy. In order to calculate the resonance energy in these saturated molecules, we may apply the simple LCAO MO treatment of conjugated molecules[10]. The following parameters are adopted in this calculation:

Coulomb integrals:

$\alpha_{H_3} = \alpha_0 - 0.3\,\beta_0$, $\alpha_{H_2} = \alpha_0 - 0.5\,\beta_0$,

Resonance integrals:

$\beta_{H_3 \equiv C} = 2.5\beta_0$, $\beta_{H_2 = C} = 2.0\beta_0$,

$\beta_{C-C} = 0.88\beta_0$ (in a C–C bond adjacent to the central C-atom),

$\beta_{C-C} = 0.50\beta_0$ (in other C–C bonds),

where α_0 and β_0 are the Coulomb and resonance integrals in benzene, respectively. The latter two parameters β_{C-C} are evaluated by considering the degree of localization in hyperconjugation[11].

The result of calculation with respect to several alkyl ions is indicated in Table VI. In S_N1 reaction the rate decreases along the series methyl < ethyl < higher *n*-alkyl < *i*-propyl < *t*-butyl, and the extremely large stability of the

7) M. J. S. Dewar, "Electronic Theory of Organic Chemistry," Clarendon Press, Oxford (1949), p. 64.
8) A. G. Evans, *Trans. Faraday Soc.*, **42**, 719 (1946).
9) J. L. Franklin, ibid., **48**, 443 (1952).
10) N. Muller and R. S. Mulliken, *J. Am. Chem. Soc.*, **80**, 3489 (1958).
11) Y. I'haya, This Bulletin, **28**, 369, 376 (1955).
12) A. Stretwieser, *Chem. Revs.*, **56**, 570 (1956).

tertiary carbonium ion was also reflected in the solvolysis rates of alkyl halides[12]. These facts are explained by our results.

Faculty of Engineering
Kyoto University
Sakyo-ku, Kyoto

Investigation on Mixed Complexes. IV. The Formation of a Mixed Complex by Reaction of Tetracyanoniccolate(II) with Bis-(dithiooxalato)-niccolate(II)

By Sigeo KIDA

(Received August 31, 1959)

In the preceding paper of this series[1], it was demonstrated that the chlorocyano- and bromocyano- platinum(II) complex could not be formed in solution. However, one can not say that the formation of any mixed cyano complexes must be ruled out. Since it is commonly known that bis-(dithiooxalato)-niccolate(II) and tetracyanoniccolate(II) ions have the planar configuration, one should expect that a reaction between these reactants in solution might produce a complex containing both cyano and dithiooxalato as ligand. The purpose of the present paper is to present spectrophotometric evidence for the formation of a mixed cyano complex of the type, $[Ni(C_2S_2O_2)(CN)_2]^{2-}$, in solution.

Equivalent quantities of potassium bis-(dithiooxalato)-niccolate(II) and potassium tetracyanoniccolate(II) were mixed in solution. At first, the dark red color of the bis-(dithiooxalato)-niccolate(II) ion was predominant in the solution, then the color slowly changed to orange, the reaction being complete in two hours at room temperature (15°C). This change was followed spectrophotometrically, and the results are shown in Fig. 1. The intensity of the absorption band at 501 mμ of the bis-(dithiooxalato)-niccolate(II) ion decreases remarkably with time, while a new band appears at 380 mμ, the intensity of which increases in nearly reciprocal proporation to that of the band at 541 mμ. This implies that a reaction such as,

$$[Ni(C_2S_2O_2)_2]^{2-} + [Ni(CN)_4]^{2-} \rightleftarrows 2[Ni(C_2S_2O_2)(CN)_2]^{2-},$$

might be taking place in solution. The reasons are: 1) the blue shift of the band from 501 mμ to 380 mμ is indicative of the replacement of the dithiooxalate ion by cyanide ions, since the former is placed behind the latter in the spectrochemical series[2], and 2) the frequency of the new band is nearly equal to the mean of frequencies of the bands of the parent complexes, $[Ni(C_2S_2O_2)_2]^{2-}$ and $[Ni(CN)_4]^{2-}$, as shown in Table I. This is the relationship which can always be seen between the mixed complex and its parent complexes[3,4].

Previously[5], polarographic measurements were carried out on a mixed solution of potassium bis-(dithiooxalato)-niccolate(II) by Maki, Kida and Tsuchida, and it was reported that a new compound, perhaps the mixed complex, $[Ni(C_2S_2O_2)(CN)_2]^{2-}$, must be formed in the solution. Thus, the present result is not in consistent with the previous one.

Similar spectrophotometric observations were made also on the bis-dithiooxalato- and the tetracyano- platinum(II) and palladium(II) complexes. Unfortunately, no indication of the expected reactions, such as

$$[Pt(C_2S_2O_2)_2]^{2-} + [Pt(CN)_4]^{2-} \rightleftarrows 2[Pt(C_2S_2O_2)(CN)_2]^{2-} \quad \text{and}$$

$$[Pd(C_2S_2O_2)_2]^{2-} + [Pd(CN)_4]^{2-} \rightleftarrows 2[Pd(C_2S_2O_2)(CN)_2]^{2-},$$

could be detected. At present, however, it is quite uncertain whether this is due to the total inertness of these parent complexes or to the extreme instability of the ligand-mixed complexes.

Therefore, another attempt was made as follows: in order to obtain some information on the possible reaction between the bis-(dithiooxalato)-platinate(II) ion and the cyanide ion, a potassium cyanide solution and a potassium

1) Part III of this series; S. Kida, This Bulletin, **33**, 587 (1960).

2) Y. Shimura and R. Tsuchida, This Bulletin, **29**, 311 (1956); R. Tsuchida, ibid., **13**, 388 (1938).

3) S. Kida, ibid., **29**, 805 (1956).

4) Y. Shimura and R. Tsuchida, ibid., **28**, 572 (1955).

5) N. Maki, S. Kida and R. Tsuchida, ibid., **32**, 573 (1959).

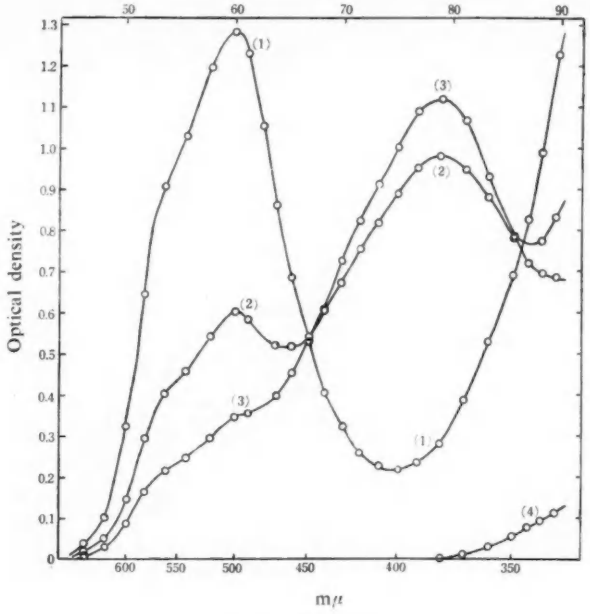


Fig. 1.

(1) $1/3\times10^{-3}$ F $K_2[Ni(C_2S_2O_2)_2]$ solution
(2) $1/3\times10^{-3}$ F $K_2[Ni(C_2S_2O_2)_2]$ and $1/3\times10^{-3}$ F $K_2[Ni(CN)_4]$ solution, (1/2 hr. after mixing)
(3) The same as solution (2), (2 hr. after mixing)
(4) $1/3\times10^{-3}$ F $K_2[Ni(CN)_4]$ solution

TABLE I. FREQUENCIES OF ABSORPTION BANDS OF THE COMPLEXES

$K_2[Ni(C_2S_2O_2)_2]$	$K_2[Ni(CN)_4]$	New band	$(\nu_1+\nu_2)/2$
59.7 (501 mμ)	96.9 (310 mμ)	78.9 (380 mμ)	78.3

$\nu\times10^{13}$: the frequency of the absorption maximum (sec^{-1})

bis-(dithiooxalato)-platinate(II) solution were mixed in the ratio 2:1. The intensity of the absorption band due to the bis-(dithiooxalato)-platinate(II) ion at 430 mμ was appreciably reduced, while in a shorter wavelength region, the absorption showed a uniform increase in intensity. No distinct new band was formed, as shown in Fig. 2.

This uniform increase in absorption intensity in the shorter wavelength region may be regarded as being due to the decomposition of the bis-(dithiooxalato)-platinate(II) ion, since it can be demonstrated spectrophotometrically that, in an alkaline solution, this ion is subject to a slow decomposition which results in increased absorption at the shorter wavelength[6]. Accordingly, a mixed complex, such as $[Pt(C_2S_2O_2)(CN)_2]^{2-}$, is unlikely to be present in the solution, though it is better not to draw a decisive conclusion until more experimental data are obtained. On the other hand, there is no doubt that some quantity of the tetracyanoplatinate(II) ion has been produced in this solution, because the characteristic band, of this ion, at 279 mμ, is evidently found in the absorption curve 6 in Fig. 2.

Experimental

Solutions of the following concentrations were mixed, $(2/3)\times10^{-3}$ F potassium bis-(dithiooxalato)-niccolate(II) and $(2/3)\times10^{-3}$ F potassium tetracyano-niccolate(II), and the mixed solution was placed in the dark at room temperature (15°C). The reaction was almost complete in two hours. It was ascertained by extinction measurements that the decomposition of bis-(dithiooxalato)-niccolate(II) ion was small during the experiment.

6) This was verified by extinction measurements on a weakly alkaline solution of this complex.
The bis-(dithiooxalato)-palladate (II) ion so rapidly decomposed in an alkaline solution that extinction measurements of this complex could not be carried out in a solution containing potassium cyanide.

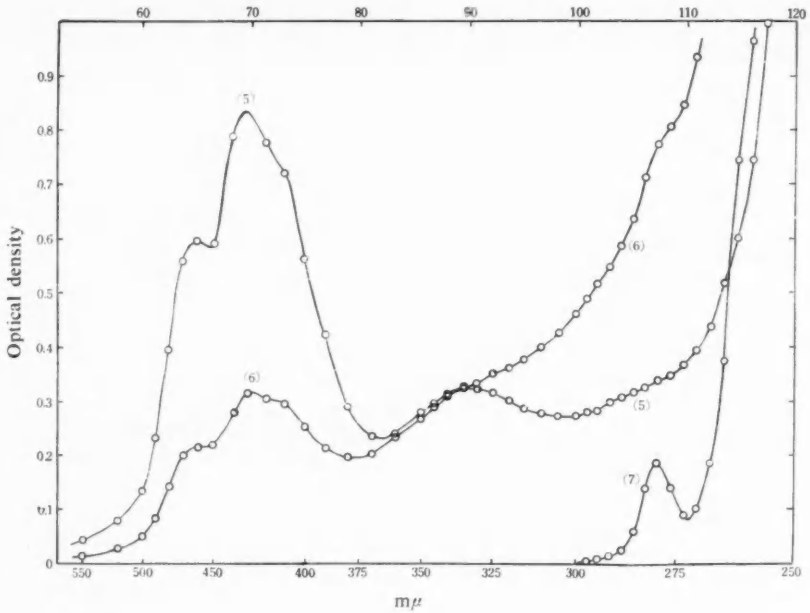


Fig. 2.

(5) 10^{-4} F $K_2[Pt(C_2S_2O_2)_2]$ solution
(6) 10^{-4} F $K_2[Pt(C_2S_2O_2)_2]$ and 2×10^{-4} KCN solution, (2 days after mixising)
(7) 10^{-4} F $K_2[Pt(CN)_4]$

Solutions of potassium cyanide (0.1 F) and of potassium bis-(dithiooxalato)-platinate(II) (0.2 F) were mixed and placed in the dark at room temperature (15°C). For extinction measurements, 5 ml. of the solution was diluted to 500 ml. The extinction measurements were carried out with a Beckman model DU spectrophotometer.

The complexes, $K_2[Ni(C_2S_2O_2)_2]$[7], $K_2Pd(C_2S_2O_2)_2]$[7], $K_2[Pt(C_2S_2O_2)_2]$[7], $K_2[Ni(CN)_4]\cdot H_2O$[8], $K_2[Pd(CN)_4]\cdot 3H_2O$[8], and $K_2[Pt(CN)_4]\cdot 3H_2O$[9], were prepared according to the literature cited.

Summary

It has been concluded, by the spectrophotometric method, that 1) potassium bis-(dithiooxalato)-niccolate(II) and potassium tetracyanoniccolate(II) react in solution to form a mixed complex, $[Ni(C_2S_2O_2)(CN)_2]^{2-}$, 2) in the case of palladium(II) and platinum(II), bis-(dithiooxalato)- and tetracyano- complexes do not react in solution, unlike the case of nickel(II), and 3) the dithiooxalate ion in the complex, $[Pt(C_2S_2O_2)_2]^{2-}$, is replaced by the cyanide ion in solution, the tetracyanoplatinate(II) ion being formed, but the formation of the mixed complex, such as $[Pt(C_2S_2O_2)(CN)_2]^{2-}$, is not certain.

The author wishes to express his heartfelt thanks to Professor R. Tsuchida for his kind instruction and advice throughout this work.

Department of Chemistry
Wakayama University
Masagocho, Wakayama

7) H. O. Jones and H. S. Tasker, *J. Chem. Soc.*, **1905**, 95.
8) W. C. Fernelius, "Inorganic Syntheses", vol. II. McGraw-Hill Book Co. Inc., 1946, pp. 243, 227.
9) G. Brauer, "Handbuch der präparativen anorganishe Chemie", F. Enke Verlag, Stuttgart, 1954, p. 1179.

Dichroism of Dyes in the Stretched PVA Sheet. IV.[1] *Absorption Spectra of Carbocyanines and Note on the Chemical Structures**

By Yoshié TANIZAKI and Hisataké ONO

(Received February 9, 1960)

In the previous paper[1], the dichroic behavior of Pinacyanol in the stretched polyvinylalcohol (PVA) sheet was discussed in detail. It was pointed out that the directions of absorptions, which were judged by an inspection of the dichroic curves observed at a high stretch ratio, are in agreement with those obtained from a quantitative analysis of the relation between the density ratio R_d and the stretch ratio R_s[1].

The purpose of this paper is to analyze the dichroism of carbocyanines by this method of inspection, to confirm the fact that the result thus obtained can be explained consistently by the conclusion derived from Pinacyanol[1], and to add some new information about the visible absorption phenomena after the inclusive discussion on the further experimental facts obtained here.

Experimental

Here were used the following four dyes, which were prepared by the known methods:

1, 1'-Diethyl-2, 2'-cyanine iodide, — m. p. 274°C (277°C in literature[2]). Found: C, 59.92; H, 5.04; N, 6.32. Calcd. for $C_{23}H_{23}N_2I$: C, 60.79; H, 5.07; N, 6.18%.

1, 1'-Diethyl-2, 2'-dicarbocyanine iodide, — m. p. 268~9°C (270°C in literature[4]).

1, 1'-Diethyl-2, 2'-tricarbocyanine iodide, — m. p. 222~3°C (225°C in literature[4]). Found: C, 61.30; H, 5.75; N, 4.80. Calcd. for $C_{29}H_{29}N_2I \cdot 2H_2O$: C, 61.30; H, 5.81; N, 4.93%.

1, 1'-Diethyl-4, 4'-cyanine iodide, — m. p. 282°C (271~3°C in literature[5]). Found: C, 60.96; H, 5.35; N, 6.37. Calcd. for $C_{23}H_{23}N_2I$: C, 60.79; H, 5.07; N, 6.18%.

Let us use the notations, $j=0(2,2')$, $j=2(2,2')$, $j=3(2,2')$ and $j=0(4,4')$, for the respective dyes described above. At the same time, the Pinacyanol used in the previous paper[1] will be denoted by $j=1(2,2')$.

The preparation of the sample sheet, the measurement, etc. followed the manner reported before[1,6]. Here, the dyeing bath of the dye which failed to dissolve in water at the necessary concentration was prepared by adding a calculated amount of water to the hot ethanol solution of the dye concentrated as much as possible.

Results

Let us number the absorption bands of each dye in the order of wavelength, for the convenience of the following description (see figures and tables).

Figs. 1 and 2 show, the interpretation of the dichroism of $j=0(4,4')$ in the near ultraviolet and in the visible regions, respectively. In Fig. 1, it is seen as a distinctive feature that the $D_{\parallel}$ and $D_{\perp}$ curves cross in the neighborhood of Band VII, that $D_{\perp}$ of Band VIII is apparently higher than D, and that D_{θ} of Band V does not change much with increasing θ.

Figs. 3 and 4 indicate the dichroism of $j=0(2,2')$ in the ultraviolet and visible regions respectively. Their special feature is that λ_{max}'s in the region longer than 300 mμ are

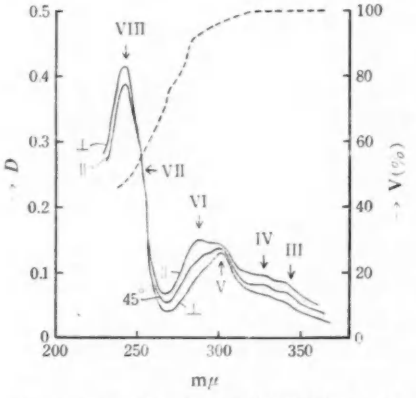


Fig. 1. The dichroism of 1,1'-diethyl-4,4'-cyanine iodide, $j=0(4,4')$, in the stretched PVA sheet. $R_s=5.7$. The broken line indicates the polarization degree curve of the polaroid used, taken from Y. Tanizaki, This Bulletin, **30**, 935 (1957).

* Presented at the Electronic Structure, Symposium, Tokyo, Sep. 1959.

1) Part III of this series, Y. Tanizaki, This Bulletin, 33, 979 (1960).

2) F. M. Hamer, *J. Chem. Soc.*, **1928**, 208.

3) F. Kimura, F. Homma and F. Kobayashi, *Japan*, **228**, 529 (1957).

4) T. Ogata and T. Shiozaki, *Bull. Iust. Phys. Chem. Res. (Tokyo)*, **13**, 511 (1934).

5) T. Ogata and S. Sakurai, *Bull. Inst. Phys. Chem. Res. (Tokyo)*, **3**, 77 (1924).

6) Y. Tanizaki, This Bulletin, **32**, 75 (1959); Y. Tanizaki and N. Ando, *J. Chem. Soc. Japan, Pure Chem. Sec. (Nippon Kagaku Zasshi)*, **78**, 542 (1957).

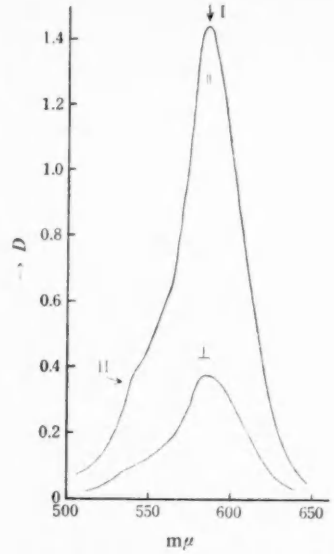


Fig. 2. The dichroism of 1,1'-diethyl-4,4'-cyanine iodide, $j=0(4,4')$, in the stretched PVA sheet. $R_s=10$.

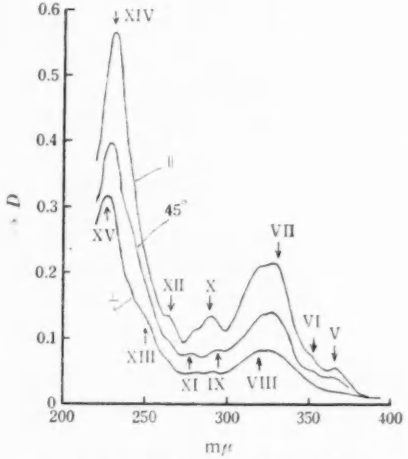


Fig. 3. The dichroism of 1,1'-diethyl-2,2'-cyanine iodide, $j=0(2,2')$, $R_s=9.0$.

almost independent of θ**, while those in the region thereunder show an apparent change. It is also an interesting fact that the shapes of absorption bands in the visible and 300～400 mμ regions resemble each other considerably (see Figs. 3 and 4), though not so closely as in the case of Pinacyanol, $j=1(2,2')$[1].

** A detailed inspection revealed a slight shift of λ_{max} with θ in the visible region. This we will discuss later.

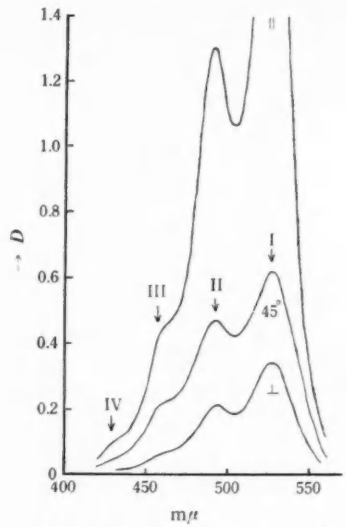


Fig. 4. The dichroism of 1,1'-diethyl-2,2'-cyanine iodide, $j=0(2,2')$, $R_s=9.0$.

Figs. 5 and 6 show the dichroism of $j=2$ $(2,2')$ at 220～500 mμ and 500～850 mμ, respectively. In this case, too, λ_{max}'s as above 400 mμ show little change with θ**, while those below 400 mμ show a clear change.

Figs. 7 and 8 show the dichroism of $j=3$ $(2,2')$. Here is found the important fact that λ_{max}'s of D_θ in Fig. 8, which shows the absorptions supposedly along the long axis of the molecule[1], apparently shift with θ. Similar behavior was also recognized in the case of $j=0$ (Fig. 4) and $j=2$ (Fig. 6), though very slightly. Band II (493 mμ) in Fig. 4, for example, shows a slight shift in λ_{max} towards shorter wavelengths with a change from $D_\perp$ to $D_\parallel$ curve. Such a tendency may be due principally to the following fact: small directional differences originally existing between the absorptions along the molecular axis come to be appreciable owing to a sharpening in the

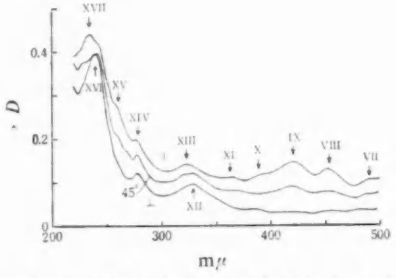


Fig. 5. The dichroism of 1,1'-diethyl-2,2'-dicarbocyanine iodide, $j=2(2,2')$. $R_s=8.1$.

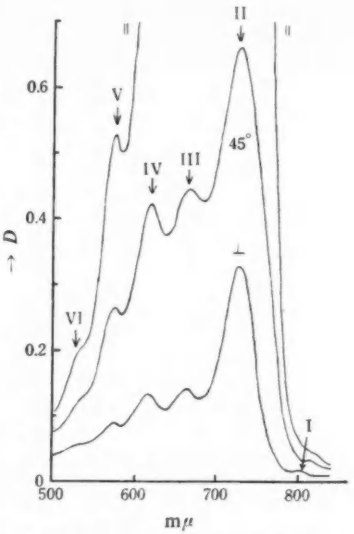


Fig. 6. The dichroism of 1,1'-diethyl-2,2'-dicarbocyanine iodide, $j=2(2,2')$. $R_s=8.1$.

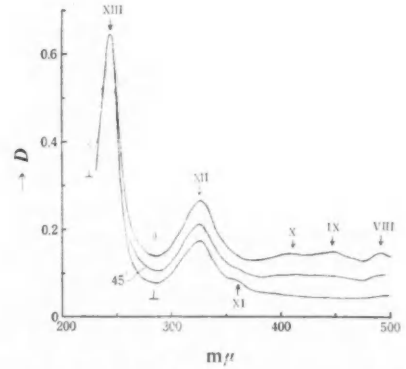


Fig. 7. The dichroism of 1,1'-diethyl-2,2'-tricarbocyanine iodide, $j=3(2,2')$. $R_s=6.1$.

distribution for the direction of the statistic orientation axis attendant on the molecule with the increase of its length. On the other hand, the directional properties of the short wavelength absorptions show a tendency to become indistinct with the increase of j. They are most indistinct in the case of $j=3$. It should be noticed that even the absorptions below 350 mμ have no change in λ_{max} with θ. Moreover, $D_{\parallel}$ and $D_{\perp}$ of Band XIII coincide with each other (Fig. 7), with the incident light of 50% polarization in this region (cf. Fig. 1).

Tables I—V involve the transition directions, the classification of band-groups, etc., as in the previous paper[1], judged from the data mentioned above. Here, table III shows again the data of Pinacyanol obtained previously[1]. Band-groups in tables have the following meaning: the absorptions in the first and the second group are both due to the transitions along the long axis of the molecule, and the absorption intensities of the second group are considerably

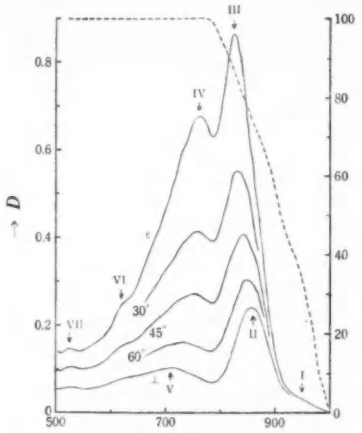


Fig. 8. The dichroism of 1,1'-diethyl-2,2'-tricarbocyanine iodide, $j=3(2,2')$. $R_s=6.9$. The broken line indicates the polarization degree curve of polaroid used, taken from Y. Tanizaki, This Bulletin, **30**, 935 (1957).

TABLE I. THE ABSORPTION BANDS OF 1,1'-DIETHYL-4,4'-CYANINE IODIDE, $j=0(4,4')$, IN THE STRETCHED PVA SHEET

Band No.	λ mμ	ν cm^{-1}	$\Delta\nu$ cm^{-1}	Direction	Group	$\Delta\tilde{\nu}$ cm^{-1}
I	585	17100	1480	l	1st	1480
II	538	18580	10640	l		
III	342	29220	1440	?	2nd	1440
IV	326	30660	2540	?		
V	301	33200	1500	$\perp(L_{\parallel})$	3rd	
VI	288	34700	4980	$\parallel(L_{\perp})$		
VII	252	39680	1620	$\parallel$?$(L_{\perp}$?)	4th	
VIII	242	41300		$\perp(L_{\parallel})$		

Note: $\Delta\nu$ and $\Delta\tilde{\nu}$ indicate the distance between the adjacent bands and the mean of the band distance among the group, respectively. The notation "l" means the longitudinal direction of the molecule, and "$\parallel$" or "$\perp$" means the direction more parallel or more normal to the l-direction, respectively, among transitions in the quinoline parts. In this table, $\parallel$ corresponds to $L_{\perp}$, the short axis of the quinoline nucleus, and $\perp$ to $L_{\parallel}$. In the following tables, $\parallel$ corresponds to $L_{\parallel}$ and $\perp$ to $L_{\perp}$.

TABLE II. THE ABSORPTION BANDS OF 1,1'-DIETHYL-2,2'-CYANINE IODIDE IN THE STRETCHED PVA SHEET

Band No.	λ mμ	ν cm^{-1}	$\Delta\nu$ cm^{-1}	Direction	Group	$\tilde{\Delta\nu}$ cm^{-1}
I	527	18960	1320	$/$		
II	493	20280	1600	$/$	1st	1450 ±150
III	457	21880	1380	$/$		
IV	430	23260	4040	$/$		
V	366	27300	1180	$/$		
VI	351	28480	950	$/$	2nd	1050 ±100
VII	330	30300	950	?		
VIII	320	31250	2650	?		
IX	295	33900	500	?		
X	290	34480	1470	∥	3rd	
XI	278	35950	1630	⊥?		
XII	266	37580	2240	∥		
XIII	251	39820	3280	⊥		
XIV	232	43100	950	∥	4th	
XV	227	44050		⊥		

Note: See the note of Table I.

TABLE III. THE ABSORPTION BANDS OF 1,1'-IODIDE DIETHYL-2,2'-CARBOCYANINE IN THE STRETCHED PVA SHEET

Band No.	λ mμ	ν cm^{-1}	$\Delta\nu$ cm^{-1}	Direction	Group	$\tilde{\Delta\nu}$ cm^{-1}
I	700	14280	1840	$/$		
II	620	16120	1260	$/$	1st	1500 ±300
III	575	17380	1480	$/$		
IV	530	18860	2410	$/$		
V	470	21270	1440	$/$?		
VI	440	22710	1670	$/$	2nd	1500 ±200
VII	410	24380	1380	$/$		
VIII	388	25760	1640	$/$		
IX	365	27400	1600	$/$?		
X	345	29000	1500	⊥		
XI	328	30500	2800	⊥	3rd	
XII	300	33300	1700	∥		
XIII	286	35000	3900	∥		
XIV	257	38900	2750	⊥		
XV	240	41650	1850	∥	4th	
XVI	230	43500		⊥		

Note: See the note of Table I. Cited from Ref. 1.

TABLE IV. THE ABSORPTION BANDS OF 1,1'-DIETHYL-2,2'-DICARBOCYANINE IODIDE IN THE STRETCHED PVA SHEET

Band No.	λ mμ	ν cm^{-1}	$\Delta\nu$ cm^{-1}	Direction	Group	$\tilde{\Delta\nu}$ cm^{-1}
I	804	12440	1400	$/$		
II	725	13790	1240	$/$		
III	665	15030	1090	$/$		
IV	620	16120	1270	$/$	1st	1300 ±200
V	575	17390	1470	$/$		
VI	530	18860	1340	$/$?		
VII	495	20200	1880	$/$?		
VIII	453	22080	1720	$/$		
IX	420	23800	1980	$/$	2nd	1800 ±100
X	388	25780	1820	?		
XI	362	27600	2880	?		
XII	328	30480	570	⊥	3rd	
XIII	322	31050	3900	∥		
XIV	278	35950	2480	⊥?		
XV	260	38430	3390	∥	4th	
XVI	239	41820	900	⊥		
XVII	234	42720		∥		

Note: See the note of Table I.

TABLE V. THE ABSORPTION BANDS OF 1,1'-DIETHYL-2,2'-TRICARBOCYANINE IODIDE IN THE STRETCHED PVA SHEET

Band No.	λ mμ	ν cm^{-1}	$\Delta\nu$ cm^{-1}	Direction	Group	$\tilde{\Delta\nu}$ cm^{-1}
I	950	10520	1110	$/$		
II	860	11630	530	$/$		ca.
III	825	12120	980	$/$	1st	1000
IV	763	13100	980	$/$		
V	710	14080	2050	$/$		
VI	620	16130	2830	$/$		
VII	527	18960	2150	$/$	2nd	ca. 2400
VIII	492	20320	2150	$/$		
IX	445	22470	1330	$/$?		
X	420	23800	3810	∥		
XI	362	27610	3150	⊥	3rd	
XII	325	30760	10370	∥		
XIII	243	41130		⊥	4th	

Note: See the note of Table I.

weak compared with those of the first one. There is a clear boundary between these two groups, as seen in their band distances, e.g. the separation of Bands IV and V in Table II is 4040 cm^{-1}. Consequently the first and the second groups can be distinguished from each other. The third and the fourth groups consist of the ∥ ⊥ character-absorptions. The third group bands are weak and the fourth strong.

According to the polarization degree curve indicated in Fig. 8, the polarization degree of incident light begins to fall precipitously at about 800 mμ, becomes about 50% at 900 mμ,

Table VI. Average wavenumber of the absorption bands in the first and the second group, difference and ratio

Dyes	$\tilde{\nu}$ cm^{-1} 1st group	2nd group	$\Delta\tilde{\nu}$ cm^{-1}	Ratio (2nd/1st)
$j=0$ (4,4')	17840	29940	12100	1.68
$j=0$ (2,2')	21100	29330	8230	1.39
$j=1$ (〃)	16660	24300	7640	1.46
$j=2$ (〃)	16260	24820	8560	1.53
$j=3$ (〃)	12290	19470	7180	1.58

and then vanishes at 1000 mμ. For this reason, it will be expected that when the absorption of 860 mμ (Band II) in Fig. 8, $j=3(2,2')$, is measured with the perfect polarized light, the band maximum will be shifted towards the longer wavelength side. Consequently, the distance between Bands I and II in Table V becomes smaller, while the distance between Bands II and III becomes greater. As a result, it will be expected that the two band distances approach the mean value of the first group in Table V.

Table VI includes the mean wave numbers of the absorption bands in the first and the second group, the difference and the ratio between these wavenumbers having been derived from Tables I—V. The ratio is approximately constant at about 1.5, except for the case of $j=0$ (4,4') dye.

Discussion

Absorptions in the Short Wavelength Region.— It has already been clarified from a detailed analysis of Pinacyanol that the absorption bands in the third and the fourth group are originated from the electronic transitions, in two directions at right angles, localized in the quinoline nucleus parts[1]. If this is the case with these absorptions, and if both (4,4') and (2,2') dye of $j=0$ have the planar structures indicated in Figs. 9(a) and (b) as respective possible ones, the dichroism similar to Pinacyanol should be expected for the latter rather than for the former. Nevertheless, according to Figs. 1 and 3, the distinct differences in direction among l, $L_{\parallel}$ and $L_{\perp}$ indicated in Fig. 9 are obviously recognized in (4,4') but not in (2,2') dye. This may be due to the steric hindrance; these molecules ($j=0$), as is well known, can not keep the planar structure. Both quinoline planes are inclined to each other, and this may be contributed mostly by the twisting of the central methine bonds. In Fig. 9(a), therefore, two $L_{\parallel}$ directions through the twisting of the methine bonds approach to being parallel with each other and normal to the l-direction. But the relative directions of $L_{\perp}$ are not altered by the twisting just mentioned. On the other hand, in Fig. 9(b), the twisting of methine bonds makes two $L_{\parallel}$'s and two $L_{\perp}$'s oblique to each other, respectively, and it may possibly occur that their oblique angles against l-direction become the same in their degree. Owing to this situation, (4,4') dye might reveal such dichroic behavior as that shown in for Fig. 1: instance the 288 mμ band (VI) has the similar tendency as the absorptions parallel with l-direction (Fig. 2), and the 301 mμ (V) and the 242 mμ band (VIII) have a tendency to be considered almost normal to l-direction. As for (2,2') dye (Fig. 9(b)), on the other hand, on account of certain resultant inclinations of $L_{\perp}$ and $L_{\parallel}$ against l-direction produced by steric hindrance, it has the dichroism indistinctive in $\parallel$ $\perp$ character as shown in Fig. 3. In contrast with this, in case of $j=1(2,2')$ (Pinacyanol) in which a steric hindrance is considered to be not so great as in $j=0$, the dichroism similar to that in Fig. 1 in appearance was observed with distinct difference in direction among $L_{\parallel}$, $L_{\perp}$ and l. This dichroism was already illustrated in the previous paper[1]. Thus, at least concerning $j=0$ dyes, the dichroism is reasonably understood by regarding the third and the fourth group absorptions as being caused by the transition at right angles to one another.

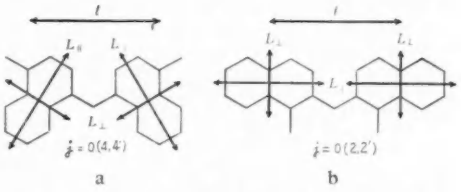


Fig. 9. Molecular skeleton diagrams of (a) 1,1'-diethyl-4,4'-cyanine cation, $j=0$ (4,4'), and (b) 1,1'-diethyl-2,2'-cyanine cation, $j=0(2,2')$.

According to the conventional consideration, the dye molecule becomes more planar with an increase in the number of methine bond, namely j. From this point of view, the $\parallel$ and $\perp$ character of absorptions of the third and fourth groups ought to be much more manifest with the increase of j. As for the (2,2') dyes, however, there is found a tendency to the contrary: the increase of j obscures the $\parallel$ $\perp$ character in dichroism. This character is observed typically in the case of Pinacyanol[1] with small j. This fact suggests that the whole strain angle of the molecule in the case of $j>1$ results from the successive summing up of the twisting angle peculiar to one methine bond. That this consideration is correct will

also be supported by the discussion in the following section.

Absorptions in the Long Wavelength Region.— It was considered previously that the second group absorptions correspond to an overtone of those of the first group[1]. One of the reasons for such consideration is also found in the dyes used here. That is to say, the second group resembles generally the first group in shape and the former is much weaker in intensity than the latter. Moreover, the ratios of their transition energies lie between 1.4 and 1.7 This fact can be interpreted in quite the same way by *second order absorption* theory on the absorptions with the same direction, proposed by Lewis et al.[7]. For this reason, there seems now no room for doubt that the absorptions of the first group and those of the second group are associated with each other as the first and the second order absorption, respectively.

It is clear from many experimental facts that the absorption phenomena of the first group of Pinacyanol is subject to remarkable change on account of its complex formation[8]. When these absorptions correspond to particular ones of respective complexes, the phenomena could be interpreted by the Davidov splitting. If so, in a case when a band in the first group is split by the formation of a complex, a corresponding absorption band in the second group with the same direction should be split in the same way. The fact that the absorptions of the first and the second group are distributed in respective groups, as shown in the tables, may suggest the very occurrence of the Davidov effect***. According to tables I—V, however, the separation of bands in the first and the second group does not change with the same tendency but with different behaviors with an increase of j (see Fig. 10) : when $j=0$, the band distance in the first group is greater than that in the second, while for $j \geqq 1$ the former becomes smaller and the latter greater with the increase of j. These relations are illustrated in Fig. 10. These obvious phenomena do not seem to be explained simply by the consideration of the Davidov effect. One possible interpretation may be given, as below, taking into account the vibrational states whose contribution might be induced by the complex formation.

Interatomic distances, usually, are lengthened by excitation to the upper state unless the ground state is forced to be strained by a particular steric hindrance. In case the molecule is already hindered in the ground state,

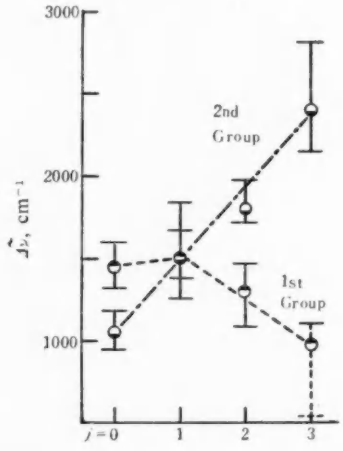


Fig. 10. Relation between $\Delta\tilde{\nu}$ and the number j of the 1st and the 2nd group absorption.

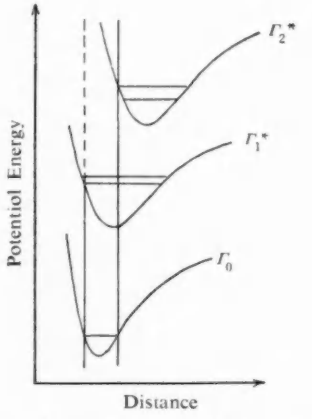


Fig. 11. Schematic potential curves for $j>1$. Γ_0 for the ground state, Γ_1^* and Γ_2^* for the first and the second excited state, respectively. The minimum parts of the respective curves shift to the right hand successively in the order of Γ_0, Γ_1^* and Γ_2^*.

7) G. N. Lewis and J. Bigeleisen, *J. Am. Chem. Soc.*, **65**, 2107 (1943).

8) I. M. Klotz, *Chem. Revs.*, **41**, 373 (1947); H. B. Klevens, *J. Phys. Chem.*, **51**, 1143(1947); R. C. Merrill, R. W. Spencer and R. Getty, *J. Am. Chem. Soc.*, **70**, 2460 (1948); R. C. Merrill and R. W. Spencer, ibid., **70**, 3683 (1948); **72**, 2894 (1950).

*** Davidov splitting of Pinacyanol under a certain condition was found for the electronic transition to the first excited state but not decidedly for the second excited state. In this case the absorption spectrum in acetone was thought to be due to the monomer, though it indicated the α and β bands. See, J. W. Weigl, *J. Chem. Phys.*, **24**, 364 (1956).

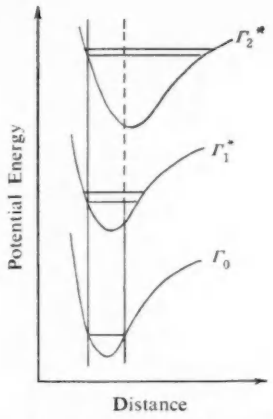


Fig. 12. Schematic potential curves for $j=0$. In this case, the minimum parts of the Γ_0 and Γ_1^* curve do not shift, but only that of the Γ_2^* curve deviates to the right side.

it may happen that the interatomic distances are not changed by excitation. Relative configurations of the potential curves in such cases would be represented schematically by Figs. 11 and 12. Fig. 11 corresponds to $j>1$, and Fig. 12 to $j=0$. In the case of Fig. 11, the minimum positions of the ground state (Γ_0), the first (Γ_1^*) and the second excited state curve (Γ_2^*) may shift towards the right hand, usually in this order. Then, if the configurational relations of Γ_1^* and Γ_2^* to Γ_0 are such as shown in Fig. 11, the band distances of the first group become smaller than those of the second group, for the higher the vibrational states are along the curve the narrower the space between the two adjacent states becomes. On the other hand, in Fig. 12 the band distances of the first group would become greater than those of the second group, if the correlation of Γ_0 and Γ_2^* curves is like that of Γ_0 and Γ_1^* curves in Fig. 11, because there is no shift between the minimum positions of Γ_0 and Γ_1^* curves as presumed above. Thus, we have one plausible interpretation of the experimental facts, according to the above consideration. At the present stage, however, there is no theoretical base for the diagrams in Figs. 11 and 12. Nevertheless, it has an instructive meaning that the phenomena shown in Fig. 10 can be explained qualitatively only by admitting such a possibility as mentioned above for the potential curves of molecules in a complex formed.

It has already been described that the absorption bands of the first group have the tendency to reveal their directional difference in dichroism with an increase of j. This does not seem to agree with conventional consideration that the increase of j makes the molecules approach to planar configurations. The reason is as follows: the better the planarity of the molecule, the better should be the parallelism of any vector in the l-direction, regardless of terminated points on both quinoline planes of a transition vector, on account of the symmetry of the molecule as seen from the structure in Fig. 9(b). That is to say, the larger the number j, the smaller must become the directional difference in dichroism. This is just opposite to the experimental facts. Thus, applying over again the consideration that the peculiar twisting angle per methine bond is added successively with an increase of j as mentioned in the last section, it stands to reason that the angles against l-direction differ in respective absorptions, so far as j is not so large and the terminals of the different transition vectors do not agree with one another; and further taking into account the sharpening of the orientation axis of the molecule, the behavior in dichroism of the first group-bands can be reasonably understood.

Summary

1. Dichroism of the electronic spectra over 220 mμ was measured for 1, 1′-diethyl-2, 2′-cyanine, -dicarbocyanine, and -tricarbocyanine iodide, and -4, 4′-cyanine iodide, respectively, in the stretched polyvinylalcohol sheets under the stretch ratio of about 6~10.

2. The correspondence of the transition directions of absorptions and the chemical structures, and the classification of the absorption bands were made into tables.

3. All the dichroic behaviors of the absorption bands could be interpreted by the transitions along the longitudinal axis of the molecule and by those, localized in the quinoline nuclei, of two directions at right angles.

4. It was confirmed that the second group absorptions correspond to the overtone of the first group ones. It was shown by the schematical potential curves that the systematic changes in the band distances in these groups with the variation of the methine bond number may be plausibly explained in connection with the vibrational states.

5. It was pointed out that a reasonable understanding for certain experimental facts in dichroism is given by regarding the molecular structures as follows; 1, 1′-diethyl-2, 2′-cyanine and -4, 4′-cyanine are far from planar because of their particular great steric hindrance, and the other cyanines are strained by the succes-

sive addition of a peculiar twisting angle per methine bond.

The authors wish to express their thanks to Professor N. Ando and Professor I. Tanaka of Tokyo Institute of Technology for their encouragement throughout this work, to Mr. Y. Mori of this Institute for his useful discussion and helpful advice, and also to Mr. Y. Hamada of the Research Laboratory of Fuji Photo Film Co., Ltd. for his help in the preparation of the dyes used.

Laboratory of Physical Chemistry
Tokyo Institute of Technology
Meguro-ku, Tokyo (Y. T.)

Research Laboratory
Fuji Film Co., Ltd.
Minami-Ashigaramachi
Kanagawa-Ken (H. O.)

A Molecular Orbital Treatment of Phosphate Bonds of Biochemical Interest. I. Simple LCAO MO Treatment

By Kenichi FUKUI, Keiji MOROKUMA and Chikayoshi NAGATA

(Received January 18, 1960)

It is well known that adenosine triphosphate (ATP) is used as a catalyst in varions kinds of in vivo chemical processes. Kalckar[1] and Lipman[2] for the first time pointed out that the energy change of hydrolysis of phosphate P–O bond of ATP was much greater than that of ordinary phosphate bonds and that this energy liberated by hydrolysis might promote endergonic reactions. Thus P–O bonds of ATP were named "high energy phosphate bonds" characterized by the sign "P~O". Further, adenosine diphosphate (ADP), phosphoenol pyruvate (PEP) and acetyl phosphate (AcP) are some of the phosphates which were believed to have high energy phosphate bonds[1-4,5]; but they can hardly be used in in vivo reactions. In this they are different from ATP[5].

Recently it has been found in calorimetrical measurements by Podolsky and Morales[6] that the enthalpy change in hydrolysis of ATP is considerably smaller than the value hitherto reported and is parallel with those of simple polyphosphates. This finding has been supported by Benzinger and Kitzinger[7], and by Burton[8]. These results suggested that there might be a revision of our understanding with respect to the role of ATP[6-8].

Kalckar[1] and Lipman[2] studied the high energy phosphate bond on the basis of the mesomeric theory. Lately Ooshika[9] assumed in their quantum-chemical discussions that the phosphate bonds of ATP were of high energy. Grabe[10] calculated the charge density of the system $-\overset{|}{\underset{\|}{C}}-O-\overset{|}{\underset{\|}{P}}-$ (O, O), and discussed at which side of the bridge oxygen the bond dissociation occurred. Gillespie et al.[11] stressed the necessity of studies on this problem from the point of view of reaction kinetics.

In this paper the authors have carried out a systematic computation on the electronic energy changes, the reactivity indexes and charge distributions of four organic phosphates, ATP, ADP, PEP and AcP, with so-called "high energy" bonds and, for comparison's sake, two phosphates lacking such bonds, adenosine monophosphate (AMP) and glycerol 1-phosphate (GLP), by a simple molecular orbital approximation. From these calculated results they have concluded that, so far as the phosphate parts only are considered, the difference, if it might exist, in the energy of

1) H. M. Kalckar, *Chem. Revs.*, **28**, 71 (1941).
2) F. Lipman, *Advances in Enzymol.*, **1**, 99 (1941).
3) P. Oesper, *Arch. Biochem.*, **27**, 255 (1950).
4) T. L. Hill and M. F. Morales, *J. Am. Chem. Soc.*, **73**, 1656 (1951).
5) E. Baldwin, "Dynamic Aspects of Biochemistry", Cambridge Univ. Press, New York, 1951.
6) R. L. Podolsky and M. F. Morales, *J. Biol. Chem.*, **218**, 945 (1956).
7) T. B. Benzinger and C. Kitzinger, *Z. Naturforsch.*, **10b**, 375 (1955).
8) K. Burton, *Nature*, **181**, 1594 (1958).
9) Y. Ooshika, *Seitai-no-kagaku*, **7**, 322 (1955).
10) B. Grabe, *Biochim. Biophys. Acta*, **30**, 560 (1958); *Archiv Fysik*, **15**, 207 (1959).
11) R. J. Gillespie, G. Maw and C. A. Vernon, *Nature*, **171**, 1147 (1953).

hydrolysis between two kinds of phosphates can not be attributed to the difference of π electronic energy change, though their reactivity is somewhat different.

Method of Calculation

The calculation is carried out in the approximation of the simple LCAO MO (linear-combination-of-atomic-orbitals molecular orbital) treatment neglecting overlap. For AcP and PEP the conjugation is spreading throughout the molecule. For ATP, ADP and AMP, on the other hand, phosphate parts only are taken into account, i.e., the adenine part, which is separated from the phosphate part by the non-conjugating ribose, is assumed not to be interacting with the phosphate part. On this assumption a few essential discussions will be given in the following section. AMP, GLP and other low energy phosphates come to have the same conjugated system in this approximation. Phosphorus atoms in the phosphates have a nearly tetrahedral, sp^3-hybridized configuration, with one electron jumping up into one of $3d$ orbitals. Strictly speaking, $2p\ \pi$ electrons on the neighboring oxygen atoms can interact not only with these d orbitals but also, though presumably only slightly, with the sp^3-hybridized orbitals[12,13]. Taking into account all of these interactions, however, complicates the problems extremely and makes it practically impossible to treat large molecules of interest. We conveniently assume, therefore, that only one hybridized d orbital on a phosphorus atom, say e.g. d^3s hybridized orbital which can afford to give a directed valence of tetrahedral symmetry, conjugates with $2p\pi$ orbitals of the neighboring oxygens in this simple treatment. Accordingly ATP, for instance, is regarded as a system consisting of 13 orbitals and 20 electrons.

The Coulomb integral of the atom or group X, that of the carbon atom attached to X and the resonance integral between that carbon atom and X are written as $\alpha+a_X\beta$, $\alpha+a_r\beta$ and $l\beta$, respectively, where α and β are the Coulomb and the resonance integrals of a carbon atom and a C–C bond in benzene, respectively. Since the precise value of a_{-O-} seems unsettled, two parallel computations are carried out for two a_{-O-} values. In Table I, together with a_X's of oxygen atoms in the phosphate part, are listed the numerical values of a_X, a_r and l adopted by us, which do not differ so much from those which were previously used by the present authors[21,22] and others[15,16]. If some oxygen atoms assigned to different a_X's should be considered to be equivalent, the mean value of their a_X's is employed; for instance, a_X of each oxygen atom in the $-PO_3^{2-}$ group is regarded as 1.0 and that in the $-PO_2^-$ group as 1.25.

Table I. Parameters used in the calculation

1. Phosphate part

X	a_X
=O	2.0
$-O^-$	0.5
–O–	0.5, 2.0

2. Others

X	a_X	a_r	l
=O	2.0	0.2	$\sqrt{2}$
–OH	0.5	0	0.6
$-CH_3$	3.0	−0.1	1

The resonance integral between a phosphorus atom and an oxygen atom, β_{P-O}, is tentatively taken $0.6\ \beta$, regardless of the bonding states of the oxygen atom (–O– type, O=type or $-O^-$ type) from a rather simple theoretical consideration on the overlap integral[13], but the numerical value itself is not essential to our results of calculation, as will be stated below. Furthermore, it is hardly possible to decide semiempirically the Coulomb integral of the phosphorus atom in the valence state in question, $\alpha_P=\alpha+a_P\beta$, because of the lack of sufficient experimental data and knowledge on the d electron phenomena. Therefore in order to avoid the unreliability of the results caused by a rather arbitrary choice of a_P, we vary the a_P value as widely as possible, as 0, −0.2, −0.4, −0.6, −1.0, −1.5, −2.0, −3.0 and $-\infty$. It should be stressed concerning this point that by means of this sliding of a_P values we can practically afford to evade not only the ambiguity of the a_P value but also the arbitrariness of the β_{P-O} value and moreover, partly, the error due to the above-stated conventional choice of the orbitals on a phosphorus atom, at least for ATP, ADP and AMP where conjugated systems are consisting of only phosphorus and oxygen atoms. In this sense the present treatment has a very general character. Judging from the ionization potential of a $3p$ electron ($I=10.98$ eV.)[10] and from Pauling's electronegativity scale ($\chi_P=2.1$), −0.6 or smaller would be appropriate for the real a_P value of a hybridized d electron.

The present authors have previously presented the frontier electron theory and have succeeded in interpreting theoretically the experience on

12) R. S. Mulliken, C. A. Rieke, D. Orloff and H. Orloff, *J. Chem. Phys.*, **17**, 1248 (1949).
13) H. H. Jaffé, *J. Phys. Chem.*, **58**, 185 (1954).
14) C. Nagata, K. Fukui, T. Yonezawa and Y. Tagashira, *Cancer Resarch*, **15**, 233 (1955).
15) H. H. Jaffé, *J. Chem. Phys.*, **20**, 279 (1952).
16) L. E. Orgel et al., *Trans. Faraday Soc.*, **47**, 113 (1951).

the chemical reactivity of π electron system[17-21]. Here we do not repeat the details of the theory but we mention only the fact that the superdelocalizability, S_r, has been known as a good intermolecular reactivity index for the substitution and the addition reactions. Also for hydrolysis of phosphates S_r is expected to remain a good measure of reactivity and is used in the present paper for discussing the ease of hydrolysis of phosphates.

Results and Discussion

The π Electronic Energy Change of Hydrolysis.—The results of calculation of the π electronic energy change of hydrolysis for two a_{-O-} values and various a_P values are listed in Table II. The listed value, $-\Delta E$, is expressed as the difference of π electronic energy between the hydrolysis products and the original reactant excepting water. For ATP, for instance, it becomes

$$-\Delta E = E(\mathrm{ADP})^{3-} + E(\mathrm{PO_4}^{3-}) - E(\mathrm{ATP})^{4-}$$

The electronic energy of water is not included in this expression because it is a constant and can not precisely be evaluated. Therefore, not the absolute value of ΔE but its difference between the two kinds of phosphate or between ATP and other phosphates has practical meaning in further discussions.

As clearly seen in Table II, ΔE of each compound does not vary so much according to a_P values. Furthermore it should be noticed that its difference between the two kinds of phosphate bonds does not exceed -0.2β for any pair of a_P and a_{-O-} values. This maximum value corresponds to ATP when a_P is zero and $a_{-O-}=2.0$. The difference for $a_P=-1.0$, a most probable value of a_P, is about $-0.12\ \beta$. Putting β as -20 kcal./mol., it corresponds to ca. 2.5 kcal./mol. If β_{P-O} is taken to be $0.8\ \beta$, it becomes ca. 3 kcal./mol. by changing the energy value of β. Thus it might be said that in the difference of the free energy change of hydrolysis between the two kinds of phosphate only 2.5 to 3 kcal./mol. at most can be ascribed to the π electronic energy.

Consequently it is concluded that, so far as phosphate parts only are taken into account, the peculiarity of ATP in several in vivo reactions cannot be attributed to the electronic energy.

Superdelocalizability for a Nucleophilic Attack and the Reactivity of Phosphate Bond.—By this time most contributions to the problem of these organic phosphate bonds of biochemical interest have been carried out from the

TABLE II. π-ELECTRONIC ENERGY CHANGE OF HYDROLYSIS

1) $a_{-O-}=0.5$.

Class	Compound	$-\Delta E$								
	$a_P=$	0	−0.2	−0.4	−0.6	−1.0	−1.5	−2.0	−3.0	−∞
Low energy	AMP, GLP etc.	$2\alpha+1.143\beta$	$2\alpha+1.168\beta$	$2\alpha+1.173\beta$	$2\alpha+1.176\beta$	$2\alpha+1.182\beta$	$2\alpha+1.187\beta$	$2\alpha+1.191\beta$	$2\alpha+1.195\beta$	$2\alpha+1.200\beta$
High energy	ATP (I)[a]	1.176	1.198	1.199	1.199	1.199	1.199	1.199	1.199	1.200
	ATP (II)[b]	0.976	0.999	1.000	1.001	1.002	1.002	1.001	1.001	1.000
	ADP	1.174	1.196	1.197	1.198	1.199	1.199	1.199	1.199	1.200
	AcP	1.181	1.201	1.201	1.201	1.200	1.197	1.195	1.192	1.175
	PEP	1.163	—	1.193	—	1.201	—	1.206	1.207	1.200

2) $a_{-O-}=2.0$.

Class	Compound	$-\Delta E$		
	$a_P=$	0	−1.0	−2.0
Low energy	AMP, GLP etc.	$2\alpha-1.654\beta$	$2\alpha-1.697\beta$	$2\alpha-1.735\beta$
High energy	ATP (I)[a]	−1.446	−1.575	−1.659
	ATP (II)[b]	−1.614	−1.763	−1.854
	ADP	−1.449	−1.576	−1.659

[a] The end "high energy" phosphate bond.
[b] The central "high energy" phosphate bond.

17) K. Fukui, T. Yonezawa and H. Shingu, *J. Chem. Phys.*, **20**, 722 (1952).
18) K. Fukui, T. Yonezawa, C. Nagata and H. Shingu, ibid., **22**, 1433 (1954).
19) K. Fukui, T. Yonezawa and C. Nagata, This Bulletin, **27**, 423 (1954).
20) K. Fukui, T. Yonezawa and C. Nagata, *J. Chem. Phys.*, **26**, 831 (1957).
21) K. Fukui, T. Yonezawa and C. Nagata, *J. Am. Chem. Soc.*, **80**, 2268 (1958).

TABLE III. SUPERDELOCALIZABILITY OF THE PHOSPHORUS ATOM FOR NUCLEOPHILIC ATTACK

1) $a_{-O-}=0.5$.

Class	Compound	$S_P^{(N)}$ $a_P=$ 0	−0.2	−0.4	−0.6	−1.0	−1.5	−2.0	−3.0	−∞
Low energy	AMP, GLP etc.	1.555	1.412	1.290	1.185	1.012	0.848	0.724	0.553	0.000
High energy	ATP (I)[a]	1.733	1.525	1.362	1.236	1.037	0.860	0.730	0.555	0.000
	ATP (II)[b]	1.810	1.576	1.401	1.259	1.049	0.865	0.732	0.556	0.000
	ADP	1.708	1.513	1.359	1.233	1.036	0.858	0.730	0.555	0.000
	AcP	1.696	—	1.371	—	1.054	—	0.728	0.561	0.000
	PEP	1.600	—	1.322	—	1.031	—	0.734	0.559	0.000

2) $a_{-O-}=2.0$.

Class	Compound	$S_P^{(N)}$ $a_P=$ 0	−1.0	−2.0
Low energy	AMP, GLP etc.	1.907	1.092	0.749
High energy	ATP (I)[a]	1.956	1.099	0.751
	ATP (II)[b]	2.511	1.191	0.778
	ADP	1.954	1.099	0.751

[a] The end "high energy" phosphate bond.
[b] The central "high energy" phosphate bond.

energetical point of view. But it seems important and interesting at this juncture for us to discuss the reactivity or the ease of hydrolysis of phosphate bonds. Hydrolysis of phosphates would be initiated by the attack of a hydroxyl anion to the phosphorus atom in vivo or of some nucleophilic group or site of tissue in vivo. So the superdelocalizability, $S_P^{(N)}$, of the phosphorus atom for a nucleophilic attack is expected to represent a theoretical index to the ease of hydrolysis. In Table III are collected the results of the calculation of superdelocalizability for a nucleophilic attack for various a_P values. The larger the $S_P^{(N)}$ is, the more reactive for a nucleophilic attack is the phosphorus atom.

As is seen in Table III, the value of $S_P^{(N)}$ of each compound is reduced markedly with the decrease of a_P values; but it is noticed that comparison of $S_P^{(N)}$ of various compounds at the same a_P value shows us some difference between the two kinds of phosphate bond, which decreases with the decrease of a_P, as 0.05 to 0.2 when $a_P=0$, 0.04 to 0.1 when $a_P=-0.4$, 0.02 to 0.04 when $a_P=-1.0$, and less than 0.01 when $a_P=-2$ and $a_P=-3$. Even larger differences of the superdelocalizability, which are realized for larger values of a_P (smaller in the absolute value), would not be sufficient to cause a casual difference of actual ease of hydrolysis. Nor could the conspicuous role of ATP be attributed to this difference, for the difference for ATP is not at all markedly greater than that for other "high energy" phosphates*. So far as phosphate parts only are considered, it would be concluded that, though some distinction can be seen in the reactivity for hydrolysis between the "high energy" bond and the ordinary one, it does not seem great enough to cause a serious difference in the role in biochemical processes.

Total π Electron Density of Oxygen Atoms. —In some biochemical problems of reactivity of conjugated molecules, effects of subsidiary atoms and bonds in the compounds were often noticed to play an essential part as well as some steric factors. A few examples of this can be seen in the carcinogenic activity of condensed aromatic hydrocarbons[14] and dimethylaminoazobenzene derivatives[22] and also in auxin activity of benzoic acid derivatives[21]. In this sense the total π electron density of oxygen atoms (q_{-O-} and q_{-O^-}) attached to a phosphorus atom may be expected to prevent electrostatically a nucleophilic reagent from approaching the phosphorus atom, where –O– and $-O^-$ stand for an oxygen atom on the phosphate bond to be broken and an oxygen anion attached to the phosphorus atom on that bond, respectively. The calculation of them has been carried out only when a_P is

* Of the two so-called high energy bonds in an ATP molecule, the inner one seems to be of very low energy (about 4 kcal. lower than the outer) from an energetic point of view, whereas the value of $S_P^{(N)}$ shows that the inner is the more reactive. It would not be possible to decide from this calculation, therefore, which bond is the easier to break.

22) C. Nagata, K. Fukui, T. Yonezawa, H. Kitano, Y. Inamoto, K. Kanai and Y. Tagashira, *Gann*, **46**, 346 (1955).

Table IV. Total π electron density and superdelocalizability of oxygen atoms for electrophilic attack[a]

1) $a_{-O-}=0.5$.

Class	Compound	$S_{-O-}^{(E)}$	q_{-O-}	$S_{-O^-}^{(E)}$	q_{-O^-}
Low energy	AMP, GLP etc.	3.055	1.868	1.797	1.914
High energy	ATP (I)[b]	2.508	1.773	1.783	1.912
	ATP (II)[c]	2.550	1.773	1.475	1.925
	ADP	2.546	1.783	1.818	1.963

2) $a_{-O-}=2.0$.

Class	Compound	$S_{-O-}^{(E)}$	q_{-O-}	$S_{-O^-}^{(E)}$	q_{-O^-}
Low energy	AMP, GLP etc.	0.952	1.952	1.743	1.991
High energy	ATP (I)[b]	0.900	1.906	1.742	1.990
	ATP (II)[c]	0.878	1.903	1.417	1.991
	ADP	0.900	1.906	1.742	1.991

[a] O^- stands for the three or the two equivalent oxygens in $-PO_3^{2-}$ or $-PO_2^-$; -O- stands for the bridge oxygen in phosphate bonds.
[b] The end "high energy" phosphate bond.
[c] The central "high energy" phosphate bond.

-1, since the choice of the value of a_P was proved in the previous paragraph not to affect the results so seriously.

On inspection of the fourth and the sixth columns of Table IV, ATP and ADP would be understood to be to a small extent preferential in hydrolysis, since the interference by the electron density in it must be smaller than in AMP.

Superdelocalizability of the Oxygen Atom for Electrophilic Attack.—In this case assuming, as was not done in the preceding paragraphs, that the initial stage of hydrolysis is the attack of proton or an electrophilic group upon the bridge oxygen atom of the phosphate bond, the superdelocalizability of the oxygen atom, $S_O^{(E)}$, for an electrophilic attack was calculated and listed in the third column of Table IV. The order of $S_{-O-}^{(E)}$ cannot explain the singularity of ATP in biochemical reactions. Nor can the order of the superdelocalizability of oxygen, $S_{-O^-}^{(E)}$, do so as clearly seen in the fifth column of Table IV.

In the foregoing discussion we found that the "high energy" phosphate bonds are not of electronically high energy, though to some extent they are more reactive, and that the conspicuous role which ATP exclusively plays in in vivo chemical reactions cannot be explained by these simple models of phosphates.

Two things must be investigated prior to the ascertainment of this conclusion. One of them is that the calculations in the present paper have been performed in the simple LCAO MO approximation: that is to say, no electron interaction has been explicitly taken into account. Interactions to be considered would occur also among $2p\pi$ electrons on a phosphorus atom. As stated in the previous section, however, since the wide sliding of a_P values from zero to minus infinity would considerably diminish the unreliability of this approximation, no serious alteration of results could be expected even in a more precise treatment.

The other is more essential. In the present paper AMP, ADP and ATP have been treated as if they were composed of only polyphosphate parts. No attention has been paid to the other conjugated part of the molecule, the adenine part, which is intervened and separated by the non-conjugated ribose group from the polyphosphate parts. This large conjugated system, however, would have some interaction with the polyphosphate part. Levendahl and James[23] found in their study on the rotatory dispersion of AMP, ADP and ATP that in ATP the last two phosphate groups and the amino group of adenine part were interacting, whereas no such interaction was noticed in the others, and suggested that this stabilized structure would an important significance in the peculiarity of the function of ATP. Szent-Györgyi[24] surmised that, as for ATP, among the 6-amino group and the 7-nitrogen of adenine parts and the two oxygens of the last two phosphate parts, the magnesium cation might intervene, through which the conjugation would spread over the molecule.

From the above calculations and discussions

23) B. H. Levendahl and J. W. James, *Biochim. Biophys. Acta*, **21**, 298 (1956).
24) A. Szent-Györgyi, "Biogenetics", Academic Press, New York, 1957.

we might say that the peculiar in vivo behavior of ATP must not be attributed to the polyphosphate part itself, but to the whole system including the polyphosphate part and the adenine part, interacting with each other directly or through a metal cation.

In the latter case the interaction of the polyphosphate part with a metal cation would considerably enhance the reactivity of its phosphorus atoms for a nucleophilic attack. Forming a weak bonding, metal cation will help them to interact. Calculation under the explicit consideration of these interactions is now going to be carried out with respect to several models for different interacting positions, extents of interaction and roles of metal cations. This will be published elsewhere.

Faculty of Engineering
Kyoto University
Sakyo-ku, Kyoto

Studies on the Configuration and Conformation of Polystyrene

By Masatami TAKEDA, Kazuyoshi IIMURA, Akira YAMADA and Yoshio IMAMURA

(Received February 13, 1960)

This work was undertaken to investigate the relationship between the chemical configuration of vinyl-type polymer determined by the process of polymerization reaction and the conformation of vinyl-type polymer determined by the physical state of a polymer.

Natta[1] showed that the polystyrene prepared by the Ziegler catalyst contains higher isotactic sequences in its chain configuration and that the crystallizability of this polystyrene is due to the helical conformation[2] of the isotactic chain which makes up the crystalline region of crystallized polystyrene.

Williams et al.[3] also found that the polystyrene prepared by the Alfin catalyst has the same tendency of crystallizability but at a lower degree.

These findings led to two different types of problems. One of them is the existence of configurational isomers, and the concept of isotactic or syndiotactic configurations. The concept of tacticity or stereoregularity is developed in this direction. The distribution of configurational isomers in each sample will be determined by the process of polymerization and solvent extraction.

The other is the existence of conformational isomers for each configurational isomer; and the distribution of conformational isomers for each configuration will be determined by the physical state of the sample and its history such as heat treatment. The dominant existence of the helical conformation of the isotactic chain in the crystalline parts of crystallized polystyrene was the only available information about polystyrene in this direction.

In the present work, the infrared spectra have been measured with four different types of polystyrenes which have been prepared by Ziegler, Natta, Alfin and Radical catalysts and by following extraction processes with cold MEK (methyl ethyl ketone). Measurements of spectra were carried out in various physical states such as quenched solid, crystallized solid, molten state and solution. The differences among spectra of the four polystyrenes in each physical state, and the differences among spectra in the physical state of each polystyrene were interpreted in the terms of configuration and conformation. Crystalline contents of samples were determined by infrared and X-ray methods. The quantitative measure of isotactic configuration has been obtained from the analysis of infrared spectra of sample solutions. The relation between the maximum content of crystallization and the measure of isotactic configuration has been studied.

Experimental

a) Sample.—Crystalline polystyrene I (CPS-I) prepared by polymerization with $AlEt_3$-$TiCl_3$ catalyst and crystalline polystyrene II (CPS-II) prepared by polymerization with $AlEt_3$-$TiCl_4$ catalyst and crystalline polystypene III (CPS-III) prepared by polymerization with Alfin Catalyst were used. Each sample of CPS was prepared from the parts insoluble in cold MEK after treating with *n*-heptane. As

1) G. Natta, *Makromol. Chem.*, **16**, 213, (1955).
2) C. W. Bunn and E. R. Howelles, *J. Polymer Sci.*, **18**, 307 (1955).
3) J. L. R. Williams, J. Vandenberche, K. R. Dunham and W. J. Dulmage, *J. Am. Chem. Soc.*, **79**, 1716 (1957).

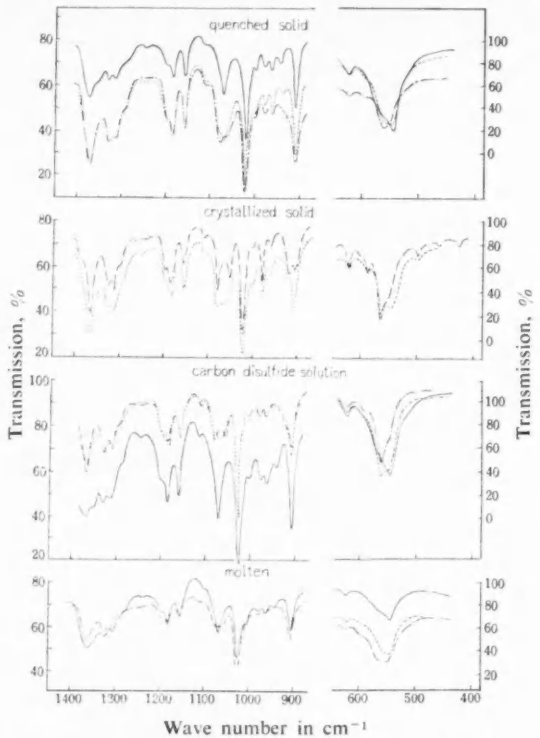


Fig. 1. Spectra of CPS-II, CPS-III and APS in various physical states.
——— APS ····· CPS-III —·—·— CPS-II

the sample of amorphous polystyrene (APS), Styron-666 of Asahi-Dow Company in Japan was used.

Film specimens for the measurement of infrared spectra were prepared by casting from carbon disulfide solution. Solution of polymers were prepared by dissolving the samples in hot carbon disulfide and the undissolved parts of CPS-I and CPS-II were filtered out.

b) Measurement of Spectra.—For measurement of the infrared absorption spectra, a Hitachi EPI-2 infrared spectrophotometer equipped with NaCl and KBr prisms was used. Spectra at the molten state of the samples were observed between 240 and 250°C by using the heating cell. Spectra of carbon disulfide solutions were observed at 2~5% concentration contained in a cell with 0.5~1.0 mm. thickness.

c) Density and X-ray Measurements.—A flotation method with NaCl solution was used for the measurement of the density at 30°C. X-ray diffraction diagrams were taken by using a Geigerflex made by Rigaku-Denki Co., Ltd.

Results and Discussion

1. Comparison of Infrared Spectra in Crystallized Solid, Quenched Solid, Solution and Molten State[4].—The infrared spectra of the four different samples have been measured in crystallized solid, quenched solid, solution and molten state. The spectra of these samples in various states are shown in Figs. 1 and 2.

In the molten state, the spectra of the four samples are almost the same. However, there

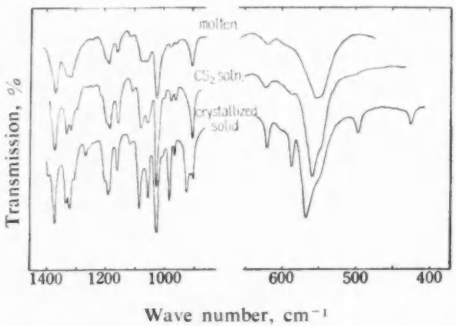


Fig. 2. Spectra of CPS-I in various physical states.

4) M. Takeda, A. Iimura, A. Yamada and Y. Imamura This Bulletin 32, 1151 (1959).

TABLE I. CLASSIFICATION OF FREQUENCIES

CPS-I			CPS-II			CPS-III			APS	Classification of the band
Cryst. solid	Soln.	Molten	Cryst. solid	Soln.	Molten	Cryst. solid	Soln.	Quenched		
1385	1385		1385							
1364	1364	1364	1364	1364	1370	1365	1367	1367	1370	B
1314	1314	1314	1314	1314	1314	1314	1314	1313	1312	B
1297	1297		1297	1297	1295	1297	1297	1295	1295	B
1258			1258							
1185	1190		1185	1190		1185	1190			B
1085	1080		1085	1080		1080	1080	1079		C_1
		1065 b			1063 b	1065	1068	1068	1070	C_2
1054	1054		1054	1057		1053	1056			C_1
984	978	978 w	984	978 w	978 w	982	978 w	979 w	980 w	A
923			923			921				A
910	908	905	910	908	905	908	908	908	905	
901			901							A
620 s	620 b	620 b	620 s	620 b	620 b	620 s	620 b	620 b	620 b	A
587	588 w, b		587	590 w, b		587	590 w, b	590 w, b		A
568			567			567				C_1
	560			560			560			C_1
		555 b			552 b					C_1
						543	544	546 b	543	C_2
499			499			499				A
			465							A
428			426			426				A

Note. (1) Only the bands which show remarkable change are listed in Table I.
(2) Remarkable change of frequencies of APS is not found with varying state.
(3) w: weak band, b: broad band, s: sharp band.

are definite differences between the band at 555 (symmetrical broad band) of CPS-I and the band at 543 cm^{-1} (asymmetrical broad band) of APS. The band at 546 cm^{-1} (asymmetrical broad band) of CPS-III appeared as the superposition of the above two bands. The rather flat and broad band at 1063 cm^{-1} of CPS-I is different from the sharp band at 1070 cm^{-1} of APS. The spectrum of molten CPS-II is similar to that of molten CPS-I.

The difference of spectra among the samples in quenched solid was very similar to the difference of those in the molten state and became larger in the crystallized solid state.

In the spectra of crystallized CPS-I and CPS-II, a new sharp band appeared at 567 cm^{-1} and the form became an asymmetric one. The broad band at 1063 cm^{-1} of molten CPS-I and CPS-II became a sharp doublet with the maxima at 1085 and 1054 cm^{-1}. In the spectrum of crystallized CPS-III, the intensity of the band increased with maximum at 567 cm^{-1}. The band around 1070 cm^{-1} of CPS-III has shown a very complicated form and probably it has three maxima at 1080, 1065 and 1053 cm^{-1}. It is also recognized that the intensity of the band at 1365, 1314, 1297 and 1185 cm^{-1} increased more in the spectra of crystallized CPS-I, CPS-II and CPS-III than in the spectra of the quenched solids as found by Tadokoro et al.[5] They suggested that these bands may be assigned to CH or CH_2 deformation vibrations and are closely related to the intramolecular interactions in the helical chain having the three fold screw axis (alternate sequences of *trans* and *gauche* conformation along the C–C chain). Other new bands with maxima at 984, 923, 901, 620, 587, 499, 465 and 426 cm^{-1} appeared in the crystallized CPS-I and CPS-II, and also the bands with maxima at 982, 921, 620, 587, 499 and 426 cm^{-1} appeared in the crystallized CPS-III. These bands are probably ascribable to the usual crystalline bands. Several heat treatments for crystallization gave no remarkable change in the spectra of APS. This fact suggests that APS does not crystallize appreciably.

The difference of spectrum became very simple in carbon disulfide solution compared to that of the crystallized solid. The bands at 1085, 1054 and 567 cm^{-1} of the crystallized CPS-I, CPS-II and CPS-III still persist in the spectra of the solution, even though slight shifts of absorbance maximum are observed. The bands at 1070 and 543 cm^{-1} of APS remain in the spectra of the solution

5) H. Tadokoro, S. Nozakura, T. Kitazawa, Y. Yasuhara and S. Murahashi, This Bulletin, 32, 313 (1959).

with no change from those of the solid. The bands at 1364, 1314, 1297 and 1185 cm^{-1} of CPS-I and CPS-II also persist in the spectra of the solution, but a slight decrease of their intensity is noted. Other bands ascribable to the crystalline band disappear or are very much weakened in the spectra of the solutions.

According to the above mentioned observation, it is possible to classify the spectra which differ from sample to sample, into three groups, A, B and C as shown in Table I.

The A group bands are characteristic in the crystallized solid and dissappear instantly or are very much weakened when the crystallized solid is dissolved in solution or is melted. These bands, which are not observed in the spectra of APS, may be assigned to the crystalline bands caused by the perturbation due to rather long range intra- and intermolecular interaction in the microcrystals of polystyrene. The assignment of A group bands to the crystalline bands makes it possible to use them for the quantitative estimation of crystalline contents of the crystallized polystyrene. The maximum crystalline content X_{max} in polystyrene might be associated with the contents of isotactic configuration in polystyrene, since only the polystyrene which has isotactic configuration can form crystalline regions with its helical conformation. This problem is studied in the 2nd part of this paper.

The B group bands are stronger in the crystallized polystyrene than in the quenched solid or the molten state, and they persist in the spectra of CPS-I, CPS-II and CPS-III even in their solutions, and are not found in that of APS. Our observations are consistent with the assignment[5] of these bands to the CH or CH_2 deformation vibration which are due to the intramolecular interaction in the helical chain conformation having a three-fold screw axis.

The C_1 and C_2 groups have shown different spectra among four samples in all physical states such as crstallized solid, quenched solid, solution and molten states. The band with maximum at 543 cm^{-1} of APS has been assigned to $\nu_1(B_2)$ vibration mode of mono-substituted benzene by Liang and Krimm[6]. According to this assignment, it can be assumed that the origin of the absorption bands which belong to the C group is the phenyl group vibration which is sensitive to the short range intramolecular interaction along the C-C main chain conformation. It is possible to subclassify the C group bands into C_1 and C_2 based on the differences of samples as shown in Table I. The C_1 group bands then become the characteristic vibrations of the C-C chain, which has the helical conformation or *trans-gauche* alternate conformation along the C-C chain from isotactic configuration. The C_2 group bands may be due to the characteristic vibration of the C-C chain which has mainly *trans* confomation along the C-C chain from syndiotactic configuration. Morton and Taylor[7] have used the new IR rating method of isotacticity based on the absorption in the 1070 cm^{-1} region of KBr disk samples. The pattern of their spectra in this region is very similar to the pettern of our spectra in the same region which are observed using the sample of crystallized film and solution.

On the basis of the above empirical classification of the characteristic bands of polystyrenes and a possible interpretation of the origin of these absorption bands, the following conclusions are tentatively drawn.

Based on the behaviors of the C class bands, the conformation of the C-C chain of a few monomer units seems to differ among the samples even in their molten states. The intensity of bands in the C_1 group will be related with the content of helical conformation of polystyrene, because its intensity becomes strong in the crystallized state. In the solution, the helical conformation of crystallizable polystyrene is partly preserved, because the persistence of the C_1 and B class bands in the spectra of CPS-I, CPS-II and CPS-III is well recognized and their intensities are stronger than these in the quenched solids. It may be suggested that the intensity ratio D560 to D543 (D560/D543) in the spectra of the solution can be used to measure the content of isotactic configuration of the same polystyrene, as well as the band around 1070 cm^{-1}. In that case the equilibrium content of helical conformation of polystyrene is determined only by its isotactic content of chain configuration, and not by its previous history like heat treatment in the solid state. The observed values of D560/D543 in solution are shown in Table II.

2. Behavior of the Crystalline Bands and the Estimation of Crystalline Content.—Based on the conclusions in part I, among several crystalline bands of the A group, the intensity of the bands at 984 and 587 cm^{-1} have been used to measure the crystalline content in various polystyrene samples. The possibility of measuring the crystalline content using the intensity of the band at 984 cm^{-1} has also been pointed out by Braun et al.[8] from their investigation on the temperature dependence of the band.

6) C. Y. Liang and S. Krimm, *J. Polymer Sci.*, **27**, 241 (1958).

7) A. A. Morton and L. D. Taylor, ibid., **38**, 7 (1959).

8) D. Braun, W. Betz and W. Kern, *Naturwissenschaften*, **46**, 344 (1959).

For avoiding the difficulty of measurement of film thickness, the band at 1030 cm^{-1} has been used as a reference, since the optical densities of this band had a linear relationship with thickness for each sample and were independent of the crystal content of samples as shown in Fig. 3.

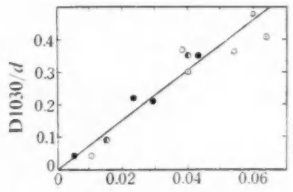


Fig. 3. Linear relation between optical density of 1030 cm^{-1} and thickness.

○: CPS-I ⊙: CPS-III
●: CPS-II ◐: APS

The measured values of D984/D1030 and D587/D1030 are listed in Table II. The linearity between both values suggests that the origin of these bands is the same crystalline band as shown in Fig. 4. The specific volumes of samples were calculated from the measured densities, and are also listed in Table II. From the specific volume, the crystalline content of polystyrene X_{sv} can be calculated by the following equation:

$$X_{sv}=(V_a-V)/(V_a-V_c) \qquad (1)$$

where, V_a is the specific volume of the amorphous part and V_c is the specific volume of the crystalline part.

If the intensities of D587 and D984 are proportional to the crystalline content, the following equation can be proved:

Table II. Numerical data for each sample

Sample	Density (30°C)	Specific volume	D587/D1030	D984/D1030	Crystalline content (%) X_i (587)	X_i (984)	X_{sv}	X_x	D560/D543 in CS_2 soln.
(CPS-I)									
Q-2	1.0529	0.950	0.042	0.082	8	5	0	11.5	
Q-2	1.0582	0.945	0.023	0.067	4	2	8	9.5	
Q-2	1.0582	0.945	0.026	0.060	5	0	8		
Q-2	1.0650	0.939	0.122	0.178	23	23	18		
Q-1	1.0652	0.939	0.122	0.172	23	22	18		
Q-1	1.0683	0.936	0.157	0.201	29	28	23		2.85
Q-1	1.0692	0.935	0.186	0.232	34	34	24	26.6	
Q-190-10	1.0742	0.931	0.204	0.248	38	37	31		
Q-205~175-310	1.0815	0.925	0.271	0.310	50	49	40		
A-220~170-480	1.0873	0.920	0.313	0.353	58	58	48		
A-220~160-480	1.0898	0.918	0.292	0.329	54	53	52		
A-220~160-750	1.0915	0.916	0.336	0.393	62	66	55	73.3*	
(CPS-II)									
Q-160-30	1.085	0.922	0.167	0.189	31	26	45	24.6	
Q-180-120	1.088	0.919	0.210	0.244	39	36	50		2.26
A-205-60	1.089	0.918	0.236	0.277	44	43	52	51.4*	
(CPS-III)									
Q-1	1.056	0.947	0.012	0.059	2	0	5		
Q-160-30	1.062	0.942	0.022	0.085	4	6	13		1.10
Q-185-120	1.068	0.936	0.078	0.132	14	14	23	16.4	
A-205-60	1.065	0.939	0.087	0.168	16	22	18	17.1*	
(APS)									
Q-1	1.052	0.951	0	0.066	0	2	—	0*	
Q-160-30	1.054	0.949	0	0.057	0	—	2		0.54
A-180-120	1.055	0.948	0	0.060	0	0	3		

Note. Q-1: Quenched in ice water after keeping the sample at 245~248°C.
Q-2: Quenched at −70°C after keeping the samples at 270°C.
Q-160-30: Means that the quenched solid was kept at 160°C during 30 minutes and after that, the temperature was lowered slowly.
A-205-60: Means that annealing was made, keeping the sample at 205°C for 60 minutes.
* These values are chosen as the maximum crystalline content X_{max}.

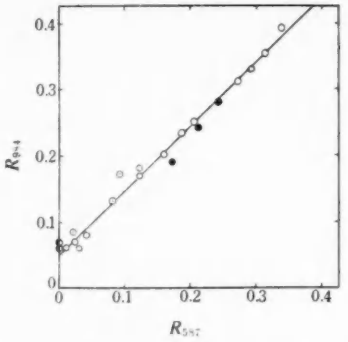


Fig. 4. Linear relation between R_{984} and R_{587}.
○: CPS-I ●: CPS-II •: CPS-III ◐: APS

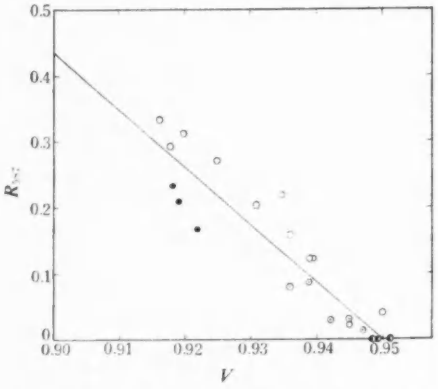


Fig. 5. Relation between R_{587} and specific volume.
○: CPS-I ●: CPS-II •: CPS-III ◐: APS

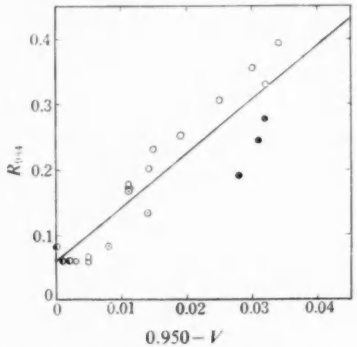


Fig. 6. Relation between R_{984} and (V_a-V).
○: CPS-I ●: CPS-II •: CPS-III ◐: APS

$$R_{587}=\mathrm{D587/D1030}=K_1(V_a-V) \quad (2)$$

$$R_{984}=\mathrm{D984/D1030}=K_2(V_a-V)+A \quad (3)$$

where, K_1 and K_2 are constant, and A is added because the band at 984 cm^{-1} has some intensity even in the amorphous state. From the linear relation between R_{587} and V, and R_{984} and (V_a-V), as shown in Figs. 5 and 6, K_1, K_2 and A were determined experimentally as follows.

$$K_1=8.729,\ K_2=8.242 \text{ and } A=0.058.$$

The value of V_a is determined, from Fig. 5, as 0.950 from the value of V at $R_{587}=0$. The value of V_c is estimated as 0.888 from the results of Natta's X-ray analysis[9].

Combining the equations (1), (2) and (3), the following relationships are obtained:

$$X_{sv}=16.13(0.950-V) \quad (4)$$

$$X_i(587)=1.85(R_{587}) \quad (5)$$

$$X_i(984)=1.96(R_{984}-0.058) \quad (6)$$

where $X_i(587)$ and $X_i(984)$ are the crystalline content by the infrared method from the 587 cm^{-1} and 984 cm^{-1} bands, respectively. The determined values of crystalline content for each sample are shown in Table II, and the agreement among X_{sv}, $X_i(587)$ and $X_i(984)$ is satisfactory considering the difference of methods. Another approximate estimation[10]

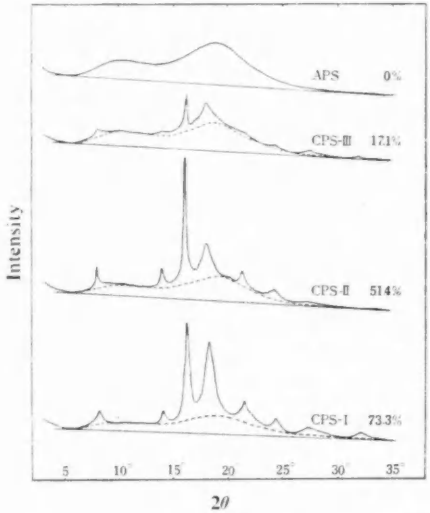


Fig. 7. X-ray diffraction diagrams of each polystyrene at maximum crystallinity and their crystalline contents.

9) G. Natta and P. Corradini, *Makromol. Chem.*, **16**, 77 (1955).
10) G. Natta et al., Atti. *accad. nazl. Lincei, Rend., Classe sci. fis., mat. e nat.*, **22**, 11 (1957).

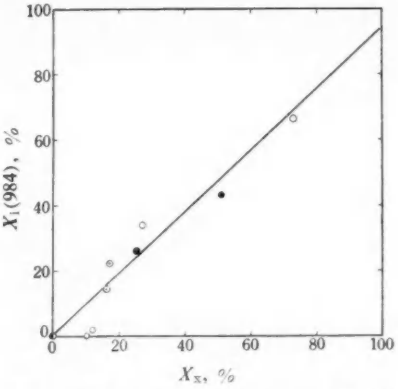


Fig. 8. Linear relation between $X_I(984)$ and X_x.

○: CPS-I ●: CPS-II ◉: CPS-III ◐: APS

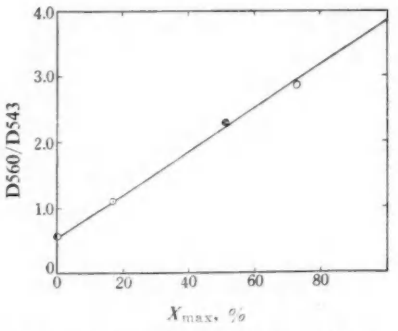


Fig. 9. Linear relation between D560/D543 and maximum crystallinity for each sample.

○: CPS-I ●: CPS-II ◉: CPS-III ◐: APS

of the crystalline content, X_x, by X-ray diffraction curve has been tried, and the values are also shown in Table II. The diffraction curve of APS in Fig. 7 has been used as the standard curve of amorphous parts. The estimated curve of amorphous parts was drawn through the point ($2\theta=12.5°$) on the experimental curve of the crystallized polystyrene making it similar to the standard curve in the direction of intensity. The intensity at $2\theta=12.5°$ was little influenced from the scattering of the crystalline parts which had the strong peaks at $2\theta=8.1°$ and $2\theta=16.2°$. The ratio of crystalline region to total area is taken as the agreement between X_x and $X_I(984)$ was satisfactory as shown in Fig. 8. The determined values of X_x for the most crystallized state of the four polystyrenes, CPS-I, CPS-II, CPS-III and APS, are taken as the maximum crystalline content, X_{max}, of each polystyrene. A strong correlation was found between X_{max} and the ratio D560/D543 in the solution, as shown in Fig. 9. This will suggest that both the maximum crystalline content, X_{max}, and D560/D543 in the solution can be used to measure the content of isotactic configuration of polystyrene from the quantitative point of view. The same treatment of the band at around 1070 cm^{-1} in the solution may be possible for the determination of the isotactic content of polystyrene, but is not considered in this paper.

We are very grateful to Mr. N. Yamazaki of the Tokyo Institute of Technology for his kind donation of the sample of CPS-I, and to Mr. S. Kondo of the Tokyo College of Science for his assistance in making the X-ray measurement.

Department of Chemistry
Tokyo College of Science
Kagurazaka, Shinjuku-ku
Tokyo

Rate of Desorption from a Heterogeneous Surface

By Yutaka KUBOKAWA

(Received February 15, 1960)

The previous studies of chemisorption of hydrogen and carbon monoxide on zinc oxide and other catalysts[1], particularly of the rates of desorption at a constant temperature, showed that, although the Elovich equation was approximately obeyed in a limited range of coverage, a slight decrease in the adsorbed amount caused a marked increase in the value of $\mathrm{d}\ln R_{\mathrm{dep}}/\mathrm{d}q$ where R_{dep} is the rate of desorption and q the adsorbed amount. A typical example is shown in Fig. 1. As was pointed out previously, such a phenomenon cannot be explained in terms of the concept that the decrease in the heat of adsorption is due to interaction between adsorbed species or to the induction effect emphasized by Boudart[2]. The concept of an a priort heterogeneity alone seems therefore to offer a satisfactory explanation of the experimental results.

Rate expressions for adsorption on such a heterogeneous surface were already derived in the previous work[3], whereby the physical meaning of $\mathrm{d}\ln R_{\mathrm{adp}}/\mathrm{d}q$ (R_{adp}=rate of adsorption) was also given. This paper presents an attempt to formulate the rate of desorption from a heterogeneous surface in a similar manner, and to explain the variation in $\mathrm{d}\ln R_{\mathrm{dep}}/\mathrm{d}q$ mentioned above on the basis of the surface heterogeneity.

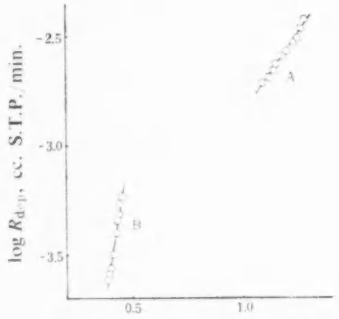


Fig. 1. Rates of desorption of carbon monoxide chemisorbed on zinc oxide. (Reproduced from "Carbon Monoxide Chemisorption on Zinc oxide" by the present author.) Temp., 192 C; The adsorbed amount at saturation is about 10 cc.; the time allowed for adsorption in run A is different from that in run B.

As described previously, the measured rate of desorption is, in most cases, unaffected by the reverse reaction, i.e., the re-adsorption. Consequently, the rate of desorption from any set of sites of the same kind on a heterogeneous surface may be

$$R=N\theta_i k\mathrm{e}^{-kt}=\alpha k\mathrm{e}^{-kt},\quad \alpha=N\theta_i$$

where θ_i is the coverage at $t=0$ and N is the distribution function of sites, both of which may usually have different values for a different set of sites. At any moment during a desorption experiment, as was found for adsorption, there is always a certain set of sites for which the rate passes through a maximum. The condition for the maximum, can be written as

$$\mathrm{d}R/\mathrm{d}k=\alpha\mathrm{e}^{-kt}(1+\beta-kt)=0,\quad \beta=\mathrm{d}\ln\alpha/\mathrm{d}\ln k$$

or
$$kt=1+\beta \tag{1}$$

Thus the maximum rate is

$$R_m=\alpha k\mathrm{e}^{-(1+\beta)} \tag{2}$$

Such a maximum rate may be assumed, as a first approximation, to be proportional to the overall rate on a similar ground to that given previously. Then, from the same argument as that given for the rate of adsorption, the following expression can be derived for the overall rate of desorption R_{dep} by using Eqs. 1 and 2:

$$\mathrm{d}\ln R_{\mathrm{dep}}/\mathrm{d}q=-(\mathrm{e}^{\beta}/RT\alpha) \tag{3}$$

In the course of desorption, k and α will vary as the set of sites for which the rate passes through a maximum is replaced by others successively. Hence, the left-hand side of Eq. 3 will generally vary as the desorption proceeds. In a limited range of coverage, however, an approximately constant value of $\mathrm{d}\ln R_{\mathrm{dep}}/\mathrm{d}q$ will be expected. The behavior of desorption such as that shown in Fig. 1 is thus in agreement with what would be expected on a heterogenous surface, although its quantitative discussion is not feasible without complete information on the site distribution function etc.

1) Y. Kubokawa, This Bulletin, **33**, 546, 550, 555, 739, 743, 747 (1960).
2) M. Boudart, *J. Am. Chem. Soc.*, **74**, 3556 (1952).
3) Y. Kubokawa, This Bulletin, **33**, 734 (1960).

The author wishes to express his sincere thanks to Professor O. Toyama for his guidance and encouragement throughout the course of this study.

Department of Applied Chemistry
College of Engineering
University of Osaka Prefecture
Sakai-shi, Osaka

Rate of Adsorption on a Heterogeneous Surface. Determination of the Distribution Function of Sites from Adsorption Rate Curves

By Yutaka KUBOKAWA

(Received February 24, 1960)

In view of the importance of surface heterogeneity in the chemisorption on oxide catalysts as revealed by desorption experiments[1], the present author has derived in a recent paper[2] the rate of adsorption on a heterogeneous surface for the cases where the distribution function of sites is variable and/or desorption can not be neglected. The rate formula thus obtained has led to an interpretation of the Elovich equation

$$dq/dt = ae^{-bq}$$

where q is the adsorbed amount and a and b are constants. Thus, explicit expressions and hence a physical meaning have been given to the constant b in the Elovich equation on the basis of the surface heterogeneity. The frequently observed variation in the constant b with pressure or temperature has been shown to be explained since the Elovich equation is usually applied only in a limited range of coverage. It is rather customary to use the integrated form of the Elovich equation for its actual application. Nevertheless, as has been pointed out in the previous paper, the constant b obtained from the integrated form of the equation, in most cases, can not be given any such physical meaning as that mentioned above, and consequently can not be related to the surface heterogeneity in a quantitative way. It has therefore been undertaken in the present work to devise a method of obtaining quantitative information on the surface heterogeneity, i. e, the distribution function of sites, directly from the adsorption rate curves.

Analysis of Adsorption Rate Curves. — As described in the previous work[2], the rate of adsorption on a heterogeneous surface can be represented by the following equation when desorption is negligibly small during the adsorption rate measurement:

$$v = Kf(p)RTNe^{-\beta} \tag{1}$$

where K and N are the rate constant of adsorption and the distribution function of sites respectively, both referring to the set of sites mainly responsible for adsorption at a given time; the value of K in Eq. 1 varies with time according to

$$Kf(p)t = 1 + \beta \tag{2}$$

where $\beta = d\ln N/d\ln K$. Thus the rate as a function of time becomes

$$v = (RTN/t)(1-\beta)e^{-\beta} \tag{3}$$

Using Eq. 1—3, the constant b of the Elovich equation, $d\ln v/dq$, is given by[3]

$$d\ln v/dq = -(1/RTN)e^{\beta} \tag{4}$$

Transposing Eq. 3,

$$d\ln t/dq = (1/RTN)(1/1+\beta)e^{\beta} \tag{5}$$

Combining Eq. 4 and 5,

$$1+\beta = -(d\ln v/dq)/(d\ln t/dq) \tag{6}$$

The right hand side of Eq. 6 is determined from the adsorption rate curves by graphical differentiation, and consequently the value of β at a given adsorbed amount, q, is obtained. Substituting β thus obtained into Eq. 1 and 4, the values of N and K are determined as a function of the adsorbed amount.

As described previously[2], the chemisorption of hydrogen on chromic oxide gel investigated by Burwell and Taylor[4] belongs to the case where desorption is negligibly small above 457°K. The data at 457°K and 1 atm. have been treated in the manner mentioned above

1) Y. Kubokawa, This Bulletin, 33, 546, 550, 555, 739, 743, 747 (1960).
2) Y. Kubokawa, This Bulletin, 33, 734 (1960).
3) The pressure term in the rate expression, $f(p)$, is assumed to be constant during adsorption.
4) R. L. Burwell and H. S. Taylor, *J. Am. Chem. Soc.*, 58, 697 (1936).

TABLE I. HYDROGEN CHEMISORPTION ON Cr_2O_3 GEL AT 457°K AND 1 atm.

q	$-\mathrm{d}\ln v/\mathrm{d}q$	$\mathrm{d}\ln t/\mathrm{d}q$	$-\beta$	$-\ln N$	$-\ln K$
3	0.163	0.495	0.67	5.66	2.02
5	0.131	0.333	0.61	5.38	2.51
7	0.092	0.244	0.62	5.04	3.06
10	0.072	0.172	0.58	4.74	3.55
12	0.064	0.142	0.55	4.60	3.79
15	0.059	0.115	0.49	4.46	4.07
20	0.053	0.092	0.43	4.30	4.45
25	0.051	0.082	0.38	4.21	4.75
30	0.047	0.075	0.37	4.12	5.07

q is in cc. (S. T. P.), v in cc./min., N in cc./cal., and K in min^{-1}.

with the results listed in Table I[5]. Here, as adsorption proceeds, the value of $\ln N$ increases with decreasing value of $\ln K$. The relationship of $\ln N$ against $\ln K$ is also obtained from a similar treatment for the data at other pressures at 457°K; the plots $\ln N - \ln K$ at different pressures at 457°K thus obtained were found to lie on the same curve, as would be expected.

By assuming that the variation in $\ln K$ arises only from that in activation energy of adsorption E, the following expression results:

$$RT \ln K = RT \ln C - E$$

where C is a constant. Plots of $\ln N$ against $\ln K$ at various temperatures combined with the above expression give those of $\ln N$ against

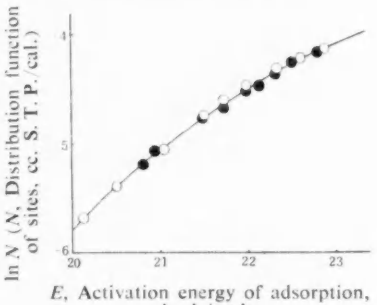


Fig. 1. Hydrogen chemisorption on Cr_2O_3 gel (Burwell and Taylor).
-○-, 457°K; -●-, 491°K; $\ln C = 20.1$.

5) $f(p) = p^{0.8}$ (p in atm.) was used, since Burwell and Taylor state that the rate of adsorption is proportional to $p^{0.8}$.

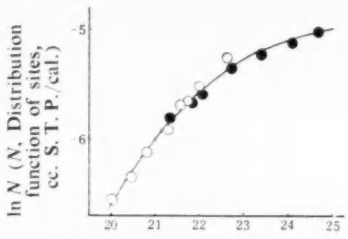


Fig. 2. Hydrogen chemisorption on $ZnO \cdot MoO_3$ (Taylor and Ogden).
-○-, 527°K; -●-, 577°K; $\ln C = 15.5$.

E all of which should lie on the same curve since N has been assumed to be a function of E alone. This requirement is fulfilled by adjusting the parameter C as shown in Fig. 1 where $\ln C = 20.1$. The figure shows that, as adsorption proceeds, the value of E increases with that of $\ln N$ from 20 to 23 kcal./mol. Values of E in this range seem to be reasonable, judging by the temperature dependence of the rate.

Fig. 2 represents the result of a similar treatment of the data of Taylor and Ogden[6] on the chemisorption of hydrogen on $ZnO \cdot MoO_3$, where desorption is negligible as described in the previous paper[1]. It is seen that the sites with a lower value of E are mainly occupied at a lower temperature.

Summary

A device to obtain the distribution function of sites from adsorption rate curves was worked out for the chemisorption for which desorption is negligible. It was then applied to the rate data of hydrogen chemisorption on chromic oxide gel and on $ZnO \cdot MoO_3$ with reasonable results.

The author wishes to express his sincere thanks to Professor O. Toyama for his guidance and encouragement throughout this work.

Department of Applied Chemistry
College of Engineering
University of Osaka Prefecture
Sakai-shi Osaka

6) G. Ogden and H. S. Taylor, *Trans. Faraday Soc.*, **30**, 1178 (1934).

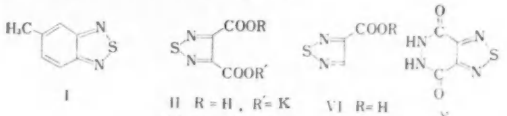
*Oxidation of 5-Methyl-2, 1, 3-benzothiadiazole with Potassium Permanganate**

By Isao SEKIKAWA

(Received February 4, 1960)

As one of the series of studies on the relation between the antitubercular activity and the molecular structure, 2, 1, 3-benzothiadiazole-5-carboxylic acid was synthesized from 5-methyl-2, 1, 3-benzothiadiazole (I) by conversion via the corresponding 5-bromomethyl- into a formyl-derivative, followed by oxidation with silver oxide.[1]

For the purpose of preparing this acid, I was oxidized with potassium permanganate at 60°C but in vain; an acid of different structure (II) was formed.[2] This new acid was finally found to be in its structure a potassium 3-carboxy-1,-2, 5-thiadiazole-4-carboxylate.



I

II R = H, R′ = K
III R = R′ = C_2H_5
V R = R′ = H

VI R = H
VII R = C_2H_5

X

Product II was an acidic potassium salt, m. p. 278°C decomp., extremely insoluble in organic solvent, with an empirical formula of $C_4HO_4N_2SK$ or $C_4H_3O_4N_2SK$ as measured by microanalysis. This salt was recrystallized unchanged from dilute hydrochloric acid and was not hydrogenated over palladium-charcoal at room temperature.

Esterification of product II yielded a colorless diethylester (III), b. p. 140~142°C (10 mmHg). The molecular formula of product III was $C_8H_{10}O_4N_2S$, as measured by microanalysis and molecular weight determination. The diethylester was converted into a diamide (IV), $C_4H_4O_2N_4S$, m. p. 234~235°C decomp., with aqueous ammonia at room temperature. Hydrolysis of the diethylester together with the diamide in boiling aqueous potassium hydroxide, followed by acidification with acetic acid, regenerated product II. From these observations the molecular formula of product II was assumed to be $C_4HO_4N_2SK$. When the diethylester was hydrolyzed with diluted hydrochloric acid, a free diacid (V) was obtained, m. p. 154~155°C, $C_4H_2O_4N_2S \cdot 2H_2O$.

On decarboxylation of diacid V in the presence of quinoline and copper powder, no definite product was obtained, but when heated at 180~190°C in vacuo, the diacid evolved carbon dioxide and was converted into a white mono acid (VI), m. p. 167°C. Acid VI was purified by sublimation, and its molecular formula was formed to be $C_3H_2O_2N_2S$, as measured by microanalysis and molecular weight determination. Mono acid VI formed

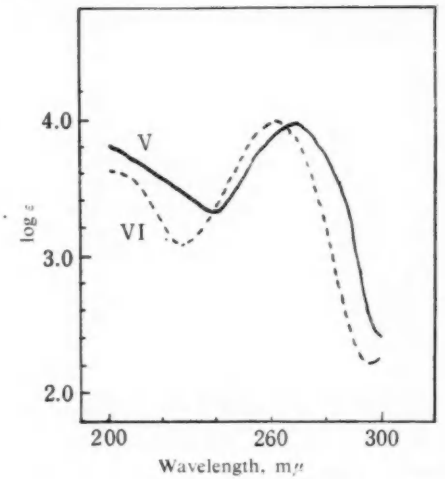


Fig. 1. Ultraviolet spectra of V and VI in water.

an ester (VII), b. p. 72°C (6 mmHg) and an acidhydrazide (VIII), m. p. 150~151°C, $C_3H_4ON_4S$. This indicates that the partial structure of mono acid VI may be $C_2HN_2S \cdot COOH$. Mono acid VI was not hydrogenated over palladium-charcoal at room temperature, being stable in boiling concentrated hydrochloric acid.

The ultraviolet spectrum (Fig. 1) of diacid V showed a pattern identical to that of acid VI. The diethylester of diacid V assigned the structure of product III, since the infrared spectra (Fig. 2) of both products VII and III showed a S–N band at 890 cm^{-1} [3].

* A part of this study was presented at the Annual Meeting of the Chemical Society of Japan, Kyoto, April 2, 1959.

1) I. Sekikawa, This Bulletin, **32**, 551 (1959).

2) I. Sekikawa, This Bulletin, **31**, 252 (1958).

3) The seven kinds of 1, 2, 5-thiadiazole compounds, showed a weak absorption band in the region of 900~880 cm^{-1}. D. H. Hadzi, *J. Chem. Soc.*, **1957**, 847; T. Momose et al., *Chem. Parm. Bull.*, **7**, 734 (1959).

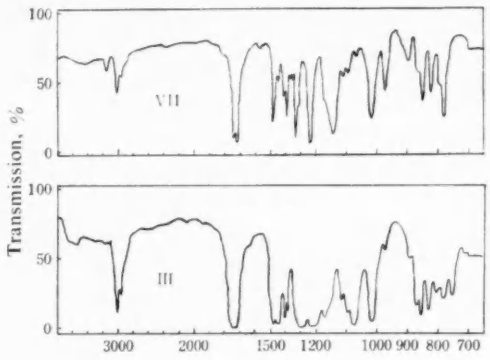


Fig. 2. Infrared spectra of VII and III.

Since the C_2N_2S-skeleton exists in sample I, it is almost certain that the white mono acid VI is 1, 2, 5,-thiadiazole-3-carboxylic acid. The presence of the 1, 2, 5-thiadiazole nucleus has been hitherto unknown.

Treatment of diethylester III with excess hydrazine hydrate in alcohol yielded almost quantitatively an orange-yellow dihydrazide (IX), m. p. 245°C decomp. with sintering from 170°C on, $C_4H_6O_2N_4S$. By hydrolysis with boiling dilute hydrochloric acid for 8 hr., one mole hydrazine was split off, from product IX and facile ring closure occurred, giving yellow 4, 5, 6, 7-tetrahydro-4, 7-dioxopyridazino-(4, 5-d)-2, 1, 3-thiadiazole (X), m. p. 333°C decomp., $C_4H_2O_2N_2S \cdot H_2O$. On the contrary, although one mole hydrazine was likewise split off, alkaline hydrolysis of product IX did not lead to product X, but to another product, which contained no sulfur atom and showed an empirical formula of $C_4H_6O_2N_4$, m. p. 331°C decomp. The determination of the structure of this compound is now in progress.

Recently, A. M. Khalexskii et al. reported that acid V was obtained by ozonolysis of 2, 1, 3-benzothiadiazole and that on the contrary alkaline permanganate oxidation gave a 2, 1, 3-thiadiazole-4, 5-dicarboxylic acid *S, S*-dioxide.[4] The oxidation of sample I to product II with potassium permanganate was a reaction similar to the oxidation of quinoline to quinolinic acid.[5]

Experimental[6]

Oxidation of I with Potassium Permanganate.—To a suspension of 10 g. of 5-methyl-2, 1, 3-benzothiadiazole in 200 ml. of water, 53 g. of potassium permanganate dissolved in 1300 ml. of water was added during 8 hr. at 60°C with vigorous stirring. After the last addition, the mixture was heated for 30 min. to complete the oxidation and then unchanged 5-methyl-2, 1, 3-benzothiadiazole was removed by steam distillation. The separated manganese dioxide was removed by filtration and washed with 500 ml. of water. The filtrates were combined, acidified with glacial acetic acid and evaporated to dryness. The residue was recrystallized from water to give 8.0 g. of white needles (II), m. p. 276～277°C (decomp.)[7]. The ultraviolet absorption bands at 270 mμ (log ε=3.94).

Found: C, 22.34; H, 0.94; N, 13.05; K, 18.58.
C, 22.29; H, 0.99; N, 12.76; K, 18.34.
C, 22.32; H, 0.71; N, 13.27.
Calcd. for $C_4HO_4N_2SK$: C, 22.64; H, 0.47; N, 13.20; K, 18.42. $C_4H_3O_4N_2SK$: C, 22.42; H, 1.41; N, 13.08; K, 18.25%.

III.—A mixture of 1.3 g. of II, 100 ml. of absolute ethanol and 1 ml. of concentrated sulfuric acid was refluxed for 24 hr. The separated potassium sulfate was filtered off and the filtrate was concentrated in vacuo. The residue was treated with 100 ml. of water and neutralized with sodium hydrogen carbonate. The mixture was extracted with 50 ml. of ether (twice), and the ether extract was dried and evaporated. The pale yellow oil was then distilled in vacuo, and a fraction boiling at 140～142°C (10 mmHg) was collected. The analytical sample distilled at 133°C (8 mmHg).

Found: C, 41.91; H, 4.61; N, 11.55; mol. wt. 235 (Rast mol. wt. determination).
Calcd. for $C_8H_{10}O_4N_2S$ (230.3): C, 41.74; H, 4.38; N, 12.17%.

IV.—To a solution of 5 g. of III in 50 ml. of ethanol, 10 ml. of aquous ammonium hydroxide (d =0.8) was added. After about 1 hr. a white precipitate separated. After the solution was kept overnight at room temperature, the white crystalline solid was collected by filtration and washed with cold ethanol; yield; 3 g. Recrystallization from ethanol yielded white needles, m. p. 234～235°C (decomp.).

Found: C, 28.21; H, 2.60; N, 32.38.
Calcd. for $C_4H_4O_2N_4S$: C, 27.90; H, 2.34; N, 32.54%.

Hydrolysis of III.—A mixture of 1 g. of III and 25 ml. of 15% hydrochloric acid was heated under reflux for 5 hr. The reaction mixture was filtered to remove a small amount of suspended solid and evaporated under reduced pressure. The residual solid was recrystallized from water to give 0.5 g. of white sandy crystals (V).

It softened at 97°C, melted at 154～155°C with effervescence and was decomposed at 160°C. The ultraviolet absorption bands at 270 mμ (log ε=3.97).

Found: C, 22.71; H, 3.05; H_2O, 16.91.
Calcd. for $C_4H_2O_4N_2S \cdot 2H_2O$; C, 22.86; H, 2.88; H_2O, 17.14%.

Pyrolysis of V.—One gram of V was placed in a sublimation tube and heated for 5 min. at 180～190°C. Under this condition the material rapidly

4) A. M. Khalexskii, V. G. Pesin and Tsin Chou, *Doklady Akad. Nauk. S. S. S. R.*, **114**, 811 (1957); *Chem. Abstr.*, **52**, 4605 (1958).
5) S. Hoogewerff und W. A. van Dorp, *Ber.*, **12**, 747 (1879).
6) All the melting points are uncorrected.
7) The m. p. of 281°C (decomp.) was recorded in the previous paper[2].

melted with evolution of carbon dioxide. The tube was then connected to a vacuum-pump and the decarboxylated product was sublimed at 190°C to give 0.5 g. of white crystals (VI), m. p. 164~165°C. The ultraviolet absorption bands at 260 mμ (log ε = 3.98).

Found: C, 27.51; H, 1.42; N, 21.50.

Calcd. for $C_3H_2O_2N_2S$: C, 27.69; H, 1.55; N, 21.53%.

Rast mol. wt. detn. gave value of 135.8, $C_3H_2O_2N_2S$ requires 130.1.

VII and VIII.—The esterification was carried out by gently refluxing for 10 hr. a mixtur of 0.9 g. of VI, 10 ml. of absolute ethanol and 0.5 ml. of concentrated sulfuric acid. After removal of the excess of ethanol, the residue was treated with water and neutralized with sodium hydrogen carbonate. The mixture was extracted with ether (twice), and the ether extract was dried, evaporated, and the residual oil was distilled to give 0.6 g. of ester (VII), b. p. 72°C (6 mmHg).

A mixture of 6.2 g. of VII, 0.2 ml. of hydrazine hydrate (80%), and 2 ml. of ethanol was heated on a water bath for 5 hr. After removing the ethanol, the solidified residue was recrystallized from ethanol, affording 0.1 g. of acid hydrazide (VIII) in the form of white needles. It softened at 140°C and melted at 150~151°C.

Found: C, 25.13; H, 3.00; N, 38.53.

Calcd. for $C_3H_4ON_4S$: C, 24.99; H, 2.80; N, 38.87%.

IX.—To a solution of 4.2 g. of diethylester III in 20 ml. of ethanol, 3.7 ml. of 80% hydrazine hydrate was added. An orange-red solid immediately separated from the reaction mixture, which was set aside overnight at room temperature and then filtered to give 3.6 g. of product. Recrystallization from water yielded yellow needles.

It softened at 150°C, reddened at 180°C and was decomposed at 248°C.

Found: C, 23.57; H, 3.20; N, 41.33.

Calcd. for $C_4H_6O_2N_6S$: C, 23.76; H, 2.99; N, 41.57%.

Hydrolysis of IX with Diluted Hydrochloric Acid.—A mixture of 2 g. of dihydrazide IX, 3 ml. of concentrated hydrochloric acid and 100 ml. of water was heated under reflux for 8 hr. Cooling and filtering yielded 1.3 g. of white crystals, which were recrystallized from water to give light yellow plates. m. p. 333°C (decomp.) (monohydrate). The ultraviolet absorption bands at 265 mμ (log ε = 4.20).

Found: C, 25.63; H, 2.35; N, 29.94; H_2O, 9.35.

Calcd. for $C_4H_2O_2N_4S \cdot H_2O$: C, 25.53; H, 2.14; N, 29.78; H_2O, 9.57%.

The filtrate from the above reaction mixture was evaporated to dryness. The residue was recrystallized from water to give white needles, m. p. 198°C (decomp.).

This material was identical with a hydrazine hydrochloride..

Hydrolysis of IV and III with Potassium Hydroxide.—A mixture of 0.5 g. of diamide IV, 0.3 g. of potassium hydroxide and 20 ml. of water was heated for 2 hr., after which time ammonia evolution had ceased. The reaction mixture was evaporated to dryness under reduced pressure. The residue was dissolved in a small amount of water, acidified with acetic acid and the precipitated solid was collected by filtration. The product was recrystallized from water to give white needles. m. p. 274°C (decomp.).

Hydrolysis of diethylester III with dilute potassium hydroxide for 1 hr., followed by isolation of the product as described above, afforded I in 60% yield.

The products from diamide and diethylester were shown to be identical with I by a microanalysis and by comparison of infrared spectra.

The author wishes to express his gratitude to Professor S. Kakimoto for his kind encouragement and advice.

The Research Institute for Tuberculosis
Hokkaido University
Sapporo

Decay of Free Radicals in γ-Irradiated Polymethyl Methacrylate

By Jiro HIGUCHI, Shōji SHIDA, Riichiro KUSAKA* and Teruo MIYAMAE

(Received February 16, 1960)

Recently, many investigations have been carried out for polymers by using the electron spin resonance (ESR) technique. Especially the polymethyl methacrylate (PMMA) is one of the most comprehensively investigated ones, and the properties of the radicals produced by γ-irradiation have been clarified fairly well[1,2]. In the present communication, a preliminary work on the decay of radicals in γ-irradiated PMMA above and below the transition point by the ESR spectrometer will be reported, while a similar work by Ohnishi and Nitta was recently carried out independently of ours only below the transition point[3].

The samples used were PMMA plates (a section 0.2×0.2 cm^2 : columnar-shaped ; Acrylite, Mitsubishi Rayon Co., Ltd.) and the irradiations were carried out 3×10^6 r. at room temperature with γ-ray from a ^{60}Co source (dose rate 2.2×10^5 r./hr.). The radicals produced were detected by the ESR spectrometer (JEP-1 E.P.R. Spectroscope ; Japan Electron Optics Laboratory Co., Ltd., X-band, bridge-type). The resonance pattern thus obtained is the same as in the case of fine powdered samples sealed in vacuo, and also similar to those by Schneider[1] and by Abraham et al.[2], and no appreciable change by the influence of oxygen was detected in the present experiment, in spite of a considerable change observed in the powdered sample exposed to air. So the experiments were carried out by using plated samples only, to avoid color centers resulting from irradiation in the glass which was used to seal the samples in vacuo.

Decay of the free radicals was investigated at temperatures between 50 and 80°C. In detection, the samples (stored in dry ice after irradiation) kept at the reaction temperature were taken out every five minutes and suddenly cooled to the room temperature by dry ice in order to interrupt the reaction rapidly, because

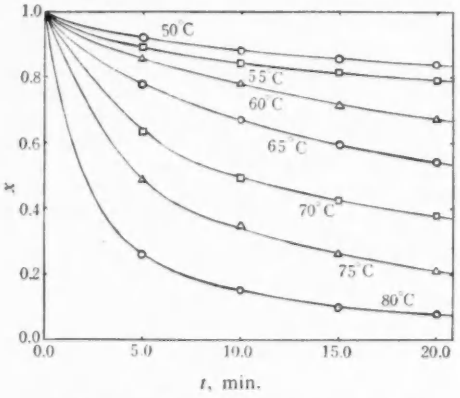


Fig. 1. Relative concentration of free radicals x aginst time t.

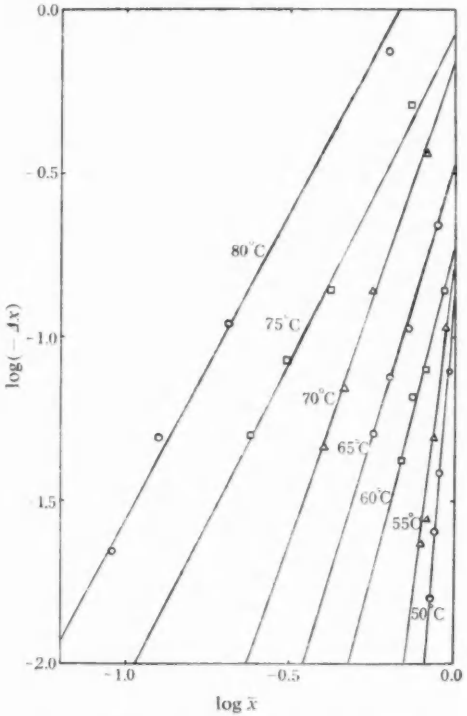


Fig. 2. Logarithm of the change of the relative concentration in five minutes $\log(-\Delta x)$ against logarithm of the mean value of the relative concentration $\log \bar{x}$.

* Present address, Yahata Iron and Steel Co., Ltd., Tokyo Research Institute.

1) E. E. Schnider, *Disc. Faraday Soc.*, **19**, 158 (1955).
2) R. J. Abraham, H. W. Melville, D. W. Overnall, and D. H. Whiffen, *Trans. Farady Soc.*, **54**, 1133 (1958).
3) S. Ohnishi and I. Nitta, *J. Polymer Sci.*, **38**, 451 (1959). In this paper, our result quoted should be corrected.

the change in the room temperature is very slow. Thus the relative concentrations in these samples were obtained by the ESR spectra at the room temperature comparing with the reference sample which was irradiated under the same condition and stored in dry ice. Under these circumstances, no appreciable change of the resonance pattern was observed.

The variation of the radical concentration so obtained is given in Fig. 1. To determine the order of the reaction, the logarithm of the change of the relative concentration in five minutes Δt, $\log(-\Delta x)$, was plotted against the logarithm of the mean value of the relative concentration, $\log \bar{x}$, as shown in Fig. 2, since the order n is obtainable from the slope by using the following relation:

$$\log(-\Delta x) = n \log \bar{x} + (\log k + \log \Delta t)$$

From Fig. 2, one can easily find the interesting feature that above 70°C the reaction is of the second order while below 70°C the apparent order is greater than the second and increases with decreasing temperature. Since the PMMA is amorphous and its transition point is considered to be about 70°C, the mechanism of the decay of the radical is considered to be different above and below the transition point taken as the border. That is, above the transition point the motion of the chain segment is rather unrestricted so that the reaction rate is proportional to the concentrations of each radical recombined, because the radicals are considered to be independent for the dose rate used, whereas below the point, the decaying process becomes complicated because of relaxation. This is also consistent with the result by Ohnishi and Nitta[3] in which below the transition point the reaction was the second order except in the beginning range (several hours). The activation energy obtained in the present work is 3×10 kcal./mol., which is also consistent with their result, although it is rather difficult to obtain accurately because of the aforementioned mechanism.

The present work is preliminary and includes many points requiring improvements as follows: (1) The error due to the fact that the samples were taken out from the reaction temperature and the ESR spectrum was observed at the room temperature. (The decay at the room temperature could not be absolutely negligible. This can be improved by the use of a cavity which can be used at various temperatures in question.) (2) A little influence of oxygen because of the experiment carried out in air. (This can be improved by sealing the sample in vacuo.) (3) The inconsistency of the experimental conditions; i.e. the γ-irradiation due to the ^{60}Co irradiation apparatus, and in the change of temperatures.

Although the present experiment encounters the above-mentioned problems, the result gives the following interesting suggestion: investigation of the decay of radicals in the γ-irradiated polymer gives information concerning the knowledge of the motion of polymers. Although it is not a direct method and the change of the mechanism bordering the transition point has already been observed in the crosslinking of polyethylene and of polyvinyl chloride[4], the present phenomenon is rather simple and favorable to the investigation of the motion of polymers. Though the present PMMA is amorphous, similar investigations can also be carried out for partially crystalline polymers. In this case, it is possible to observe separately the motion of polymers in the crystalline and amorphous regions since the resonance patterns for these regions can not always be the same because of the steric effect of the radical produced. In fact, including the case of PMMA, a more precise investigation is absolutely necessary, especially for the points aforementioned, in order to obtain more conclusive result.

The authors wish to thank Dr. M. Kurata for his helpful advice.

Tokyo Institute of Technology
Ookayama, Meguro-ku, Tokyo
(J. H. & S. S.)

Japan Electron Optics Laboratory Co., Ltd.
Kamirenjaku, Mitaka, Tokyo
(R. K. & T. M.)

4) A. N. Pravednikov, Y. S. Kan, and S. S. Medvedev, *Proc. United Nations International Conference on the Peaceful Uses of Atomic Energy*, No. 2294, 192 (1958); See, S. Okamura, *Chemistry (Kagaku)*, **5**, 7 (1960).

*Structure of Namakochrome. II**

By Toshihiko MUKAI**

(Received February 12, 1960)

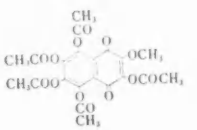
Previous paper[1] has reported that namakochrome which is a new pigment isolated from *polycheira rufescens*, should be either 2-hydroxymethyl-3, 6, 7-trihydroxy-naphthazarin (I or II) or 2-methoxy-3, 6, 7-trihydroxy-naphthazarin (III or IV). Thus, for the complete elucidation of the namakochrome structure there remains to be established the form of the functional group in 2-position of the naphthalene nucleus.

I II

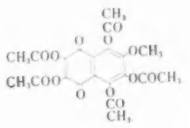
III IV

Although, in the preceding work[1], a qualitative detection of the quinone group in namakochrome by reductive acetylation has been also described, the isolation of crystalline products from the resulting solution was unsuccessful. Namakochrome-trimethyl ether was then acetylated in the presence of zinc dust, and a benzene solution of the reaction products was filtrated through an alumina column, and afforded a leuco-acetyl derivative, $C_{11}H_3O_5\cdot(OCH_3)_3(COCH_3)_4$, as colorless prisms, m. p. 226~227°C.

Now, in order to characterize the unascertained oxygen atom in 2-position of namakochrome, the 3, 5-dinitrobenzoate was prepared by the action of 3, 5-dinitrobenzoyl chloride on namakochrome-trimethyl ether. By carrying out this process in the presence of pyridine, there were obtained lemon yellow prisms, melting at 218~220°C. Its analytical composition is in agreement with a formula $C_{11}H_3O_5\cdot(OCH_3)_3[C_6H_3(NO_2)_2(COO)]_2$. The fact that namakochrome-trimethyl ether reacted with two molecular equivalents of 3, 5-dinitrobenzoyl chloride suggests, that the positions for the 3, 5-dinitrobenzoyl groups in the naphthalene nucleus should be at 5 and 8. The functional group at 2-position of the naphthalene nucleus might be a methoxyl group because of its resistance to both of the acetylation[1] and the 3, 5-dinitrobenzoylation. A qualitative detection of the methoxyl group in namakochrome was then made. In the alkoxyl microapparatus of Zeisel, namakochrome was treated with hydriodic acid as usual. The resulting vapor was passed through a bubbler which contained a mixed aqueous solution of sodium thiosulfate and cadmium sulfate, and then swept into a 10% alcoholic solution of dimethylaniline. Thus, it was possible to isolate trimethyl phenyl ammonium iodide, colorless plates, subliming at 216°C. This showed a methoxyl group to be present in namakochrome. The quantitative determination was then carried out by the same method, and namakochrome had shown an average value of 9.26% for the methoxyl group content (calcd. for III or IV: CH_3O, 11.6%). Similarly the determination was effected on a few derivatives of namakochrome described in the previous paper[1] and in this paper. A pentaacetyl-derivative[1] was found to show the methoxyl group content of 4.42% (calcd. for V and VI: CH_3O, 6.5%).

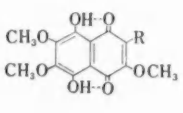


V



VI

The result for trimethyl ether of namakochrome was 37.38% of the methoxyl group content (calcd. for VII: $4CH_3O$, 40.0%, and for VIII: $3CH_3O$, 30.0%), and then for the leuco-acetyl derivative was found the methoxyl group content of 23.38% (calcd. for IX: $4CH_3O$, 25.81%, and for X: $3CH_3O$, 19.24%).

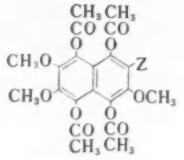


VII, R=-OCH_3,
VIII, R=-CH_2OH.

* "Pigments of Marine Animals", X. IX of this series; This Bulletin, 33, 453 (1960).

** Present address, Yoshitomi Pharmaceutical Industries, Ltd. Yoshitomi Factory, Yoshitomi-cho, Chikujo-gun, Fukuoka-ken.

1) This Bulletin, 33, 453 (1960).



IX, Z=-OCH_3,

X, Z=-CH_2OH.

From these experimental results, it is considered to be very likely that the structure of namakochrome is 2-methoxy-3, 6, 7-trihydroxynaphthazarin (beta-position in I or II : methoxyl group)***. Although many compounds which contain methoxyl groups have been known as plant components, the author has been unable to find in the literature any reports on the marine-animal pigments, especially on naphthoquinone-derivatives, which contain methoxyl groups. This must be, therefore, the first time, that the pigment containing a methoxyl group is isolated from the marine animal.

Experimental

Reductive Acetylation of Namakochrome-trimethyl Ether.—To a solution of 55 mg. of namakochrome-trimethyl ether in 1 g. of acetic anhydride and 1.5 g. of anhydrous pyridine was added 500 mg. of zinc dust, and shaken at room temperature until the solution becomes colorless. But the colorless solution soon acquired a yellow color and finally an orange yellow color. The colored reaction mixture, after setting overnight and removal of zinc dust, was poured into ice-water and extracted with several portions of benzene. The yellow benzene layer was washed with dilute hydrochloric acid solution and with water. The resulting colorless benzene solution was then dried with sodium sulfate. Removal of all solvent under the diminished pressure left 72 mg. of a pale yellow colored residue. It was purified by recrystallization from ethanol ; colorless prisms, m. p. 226~227°C.

Found : C, 55.26 ; H, 5.22 ; CH_3O-, 23.38 ; CH_3CO-, 35.27. Calcd. for $C_{22}H_{24}O_{12}$: C, 55.00 ; H, 5.01 ; CH_3O-, 25.81 ; CH_3CO-, 35.87%.

5, 8-Di-(3', 5'-dinitrobenzoyl)-namakochrome-trimethyl Ether.—40 mg. of namakochrome was dissolved in 10 ml. of hot anhydrous pyridine. To the crimson-colored solution, about 55 mg. of finely pulverized 3, 5-dinitrobenzoyl chloride was added at room temperature, and the reaction mixture was gently shaken. The color changed through yellow to orange red. After an hour, the mixture was poured into ice-water and extracted with benzene. The orange red benzene layer was washed with dilute hydrochloric acid solution, followed by dilute sodium carbonate solution and water. The orange red benzene solution was dried with sodium sulfate, and concentrated to a small volume under reduced pressure. The concentrate was filtered through a calcium carbonate column (3×15 cm.). The orange eluate was evaporated to dryness under reduced pressure, and there was obtained an orange-yellow pasty product which was recrystallized (acetone-ethanol=2 : 1) to afford 66 mg. of the 3, 5-dinitrobenzoate, lemon yellow prisms, m. p. 218~220°C.

Found : C, 48.41 ; H, 3.09 ; N, 7.70. Calcd. for $C_{29}H_{15}O_{18}N_4$: C, 48.15 ; H, 2.60 ; N, 8.02%.

I am deeply indebted to Professor Dr. Tokuichi Tsumaki for his continuing interest in this work and particularly for his suggestions. The microanalyses were carried out under Mr. Michio Shido and Miss. Sachiko Indo, to whom I extend my thanks. The expenses of the work were partially covered by a scientific Research Grant of the Ministry of Education.

Department of Chemistry
Faculty of Science
Kyushu University
Hakozaki, Fukuoka

*** That tautomerisation of this type readily occurs had been synthetically demonstrated[2] in the case of methylnaphthazarin, which was prepared (a) by condensing maleic anhydride with toluquinol and (b) by condensing citraconic anhydride with quinol, and the two products were identical.

2) A. K. Macbeth et al., *J. Chem. Soc.*, 1935, 333.

The Kinetics of Substitution Reactions Involving Metal Complexes. IV. Reaction between Lead and Zinc-Ethylenediaminetetraacetate Complexes in Aqueous Solutions

By Nobuyuki TANAKA and Kiyoko KATO*

(Received February 22, 1960)

The substitution reaction of lead ions and zinc-ethylenediaminetetraacetate (EDTA) complexes was investigated by Bril, Bril and Krumholz[1]. The rate of the reaction was followed by measuring the change of the polarographic diffusion current of lead ions with time in acetate buffer solutions using the streaming mercury electrode. The reaction was found to proceed, under the experimental conditions, through three different reaction paths simultaneously, and the rate constants of several reactions were determined.

Although Bril et al.[1] measured the reaction rate in acetate buffer solutions, they did not take consideration of the effect of acetate ions on the reaction rate. Recently, the present authors[2,3] dealt with the substitution reaction of copper ions and lead-EDTA complexes and that of copper ions and zinc-EDTA complexes, and explained the decrease of the reaction rate with increasing concentration of acetate by assuming that the substitution reaction proceeds only through the hydrated metal ions.

In this situation it seems worth while to reinvestigate the substitution reaction of lead ions and zinc-EDTA complexs from the same viewpoint as given in the previous papers[2,3].

Experimental

The preparation of the solutions of lead nitrate, zinc nitrate and disodium ethylenediaminetetraacetate was described in the previous papers[3,4].

The rate of the reaction between lead ions and zinc-EDTA complexes was followed by measuring the change of the polarographic diffusion current of lead ions with time in acetate buffer solutions. Current-time curves were obtained with a Shimadzu RP-2 pen-recording polarograph and a dropping mercury electrode. The latter was the same as described in the previous papers[2-4] and had an m value of 1.27 mg./sec. and a drop time of 3.69 sec. in an air-free 0.1 M potassium chloride solution at 25°C with open circuit.

Measurements were made at 15 and 25°C in acetate buffers of pH 4.8 to 5.7 which contained 0.1 M of sodium acetate. The ionic strength was adjusted to be 0.2 with potassium nitrate, unless otherwise stated. Gelatin was added as a maximum suppressor by 0.01% in concentration.

Results and Discussion

In the solution containing acetate ions less than 0.2 M in concentration, lead(II) forms mono- and diacetato complexes, whereas zinc(II) forms only a monoacetato complex[5]. Therefore, the overall substitution reaction of lead ions and zinc-EDTA complexes in acetate buffer solutions is represented with the equations

$$\begin{aligned} Pb^{2+} + 2\,OAc^- &\rightleftharpoons PbOAc^+ + OAc^- \\ &\rightleftharpoons Pb(OAc)_2 \end{aligned} \quad (1)$$

$$Pb^{2+} + ZnY^{2-} \rightleftharpoons PbY^{2-} + Zn^{2+} \quad (2)$$

$$Zn^{2+} + OAc^- \rightleftharpoons ZnOAc^+ \quad (3)$$

where Y^{4-} means a completely dissociated EDTA anion. Under the condition that both zinc ions and zinc-EDTA complexes are present in a large excess over lead ions, the substitution reaction 2 can be treated simply as

$$Pb^{2+} \underset{k^-}{\overset{k^+}{\rightleftharpoons}} PbY^{2-} \quad (4)$$

The rate equation for the reaction 4 is expressed as[2]

$$-\frac{d[Pb^{2+}]_{app}}{dt} = k^+[Pb^{2+}]_{app} - k^-[PbY^{2-}] \quad (5)$$

where $[Pb^{2+}]_{app}$ is the concentration of lead ions that can be determined polarographically, and given by the equation

$$\begin{aligned} [Pb^{2+}]_{app} = [Pb^{2+}] &+ [PbOAc^+] \\ &+ [Pb(OAc)_2] \end{aligned} \quad (6)$$

Integrating Eq. 5,

$$\begin{aligned} \ln\{[Pb^{2+}]_{app} - [Pb^{2+}]_{e,app}\} &= -(k^+ + k^-)t \\ &+ \ln\{[Pb^{2+}]_{o,app} - [Pb^{2+}]_{e,app}\} \end{aligned}$$

is obtained, where subscripts o and e denote

* Present addess: The Women's Department, Tokyo College of Pharmacy, 31, Uenosakuragi-cho, Taito-ku, Tokyo.

1) K. Bril, S. Bril and P. Krumholz, *J. Phys. Chem.*, **59**, 596 (1955); **60**, 251 (1956).

2) N. Tanaka and K. Kato, This Bulletin, **32**, 1376 (1959).

3) K. Kato, ibid., **33**, 600 (1960).

4) N. Tanaka, K. Kato and R. Tamamushi, ibid., **31**, 283 (1958).

5) N. Tanaka and K. Kato, This Bulletin, **33**, 417 (1960).

the initial and the equilibrium condition, respectively. At the equilibrium of the reaction, the relation

$$\frac{k^+}{k^-} = \frac{[PbY^{2-}]_e}{[Pb^{2+}]_{e,app}} \fallingdotseq K^{Pb}_{Zn,app} \frac{[ZnY^{2-}]_0}{[Zn^{2+}]_{0,app}}$$

may be satisfied, where $K^{Pb}_{Zn,app}$ represents the apparent equilibrium constant expressed as

$$K^{Pb}_{Zn,app} = \frac{[Zn^{2+}]_{e,app}[PbY^{2-}]_e}{[Pb^{2+}]_{e,app}[ZnY^{2-}]_e}$$

The values of $K^{Pb}_{Zn,app}$ were determined polarographically to be 9.0 at 15°C and 5.4 at 25°C in acetate buffer solutions which contained 0.1 M of sodium acetate.

The apparent rate constant, $k^+/[ZnY^{2-}]_0$, was obtained at varied pH's at two different concentrations of zinc ions but at the same concentrations of other species. The values obtained were plotted against the hydrogen ion concentration. The plots obtained for 15°C are given in Fig. 1, which clearly indicate linear relations existing between $k^+/[ZnY^{2-}]_0$ and $[H^+]$. Similar relations were also found at 25°C.

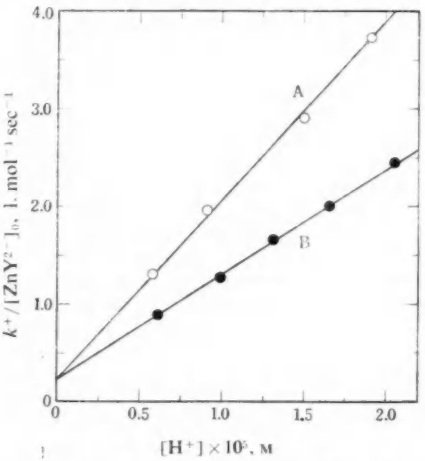


Fig. 1. $k^+/[ZnY^{2-}]_0$ as a function of the hydrogen ion concentration in acetate buffer–potassium nitrate solutions of ionic strength 0.2 at 15°C. Initial concentrations are: $[Pb^{2+}]_{0,app} = 1.60 \times 10^{-4}$ M; $[ZnY^{2-}]_0 = 5.04 \times 10^{-3}$ M; $[Zn^{2+}]_{0,app} = 5.10 \times 10^{-3}$ M (curve A) and 1.02×10^{-2} M (curve B). Concentrations of free acetate are: 0.098 M (A) and 0.096 M (B).

According to Bril, Bril and Krumholz[1], the substitution reaction between lead ions and zinc-EDTA complexes is supposed to proceed simultaneously through the following three reaction paths:

(i) $Pb^{2+} + ZnY^{2-} \rightleftharpoons PbY^{2-} + Zn^{2+}$ (7)

(ii) $ZnY^{2-} + H^+ \rightleftharpoons ZnHY^-$ (8a)

$Pb^{2+} + ZnHY^- \rightleftharpoons PbHY^- + Zn^{2+}$ (8b)

$PbHY^- \rightleftharpoons PbY^{2-} + H^+$ (8c)

(iii) $ZnY^{2-} + H^+ \rightleftharpoons Zn^{2+} + HY^{3-}$ (9a)

$Pb^{2+} + HY^{3-} \rightleftharpoons PbY^{2-} + H^+$ (9b)

In the polarographic measurement in acetate media, the concentration of lead ions that is actually measured polarogrphically is not the concentration of hydrated lead ions but the sum of the concentrations of hydrated ions and acetato complexes[2]. Therefore, if it is assumed that the substitution reaction proceeds only through the hydrated ions which are in equilibrium with the acetato complexes, the apparent rate of the forward reaction ($v^+ = k^+[Pb^{2+}]_{app}$) determined polarographically is represented by Eq. 10[2,3].

$$v^+ = k^+[Pb^{2+}]_{app} = \left\{ \frac{k_1^+ + k_2^+[H^+] + k_3^+[H^+]\dfrac{1 + K_{ZnOAc}[OAc^-]}{[Zn^{2+}]_{app}}}{1 + \sum_{n=1}^{2} K_{Pb(OAc)_n}[OAc^-]^n} \right\} \times [Pb^{2+}]_{app}[ZnY^{2-}] \quad (10)$$

In Eq. 10

$$\left.\begin{aligned} k_2^+ &= k_{II}^+ K^{II}_{ZnHY} \\ k_3^+ &= k_{III}^+ / K_{ZnY} K_{HY} \end{aligned}\right\} \quad (11)$$

where K^{II}_{ZnHY} represents the equilibrium constant of the reaction 8a, K_{ZnY}, the formation constant of ZnY^{2-} and K_{HY}, the fourth dissociation constant of ethylenediaminetetraacetic acid. $K_{Pb(OAc)_n}$ and K_{ZnOAc} represent the overall formation constants of acetatolead(II) and acetatozinc(II) complexes, respectively.

From Eq. 10, Eq. 12 is obtained, where the initial concentrations, $[Zn^{2+}]_0$ and $[ZnY^{2-}]_0$, are substituted for $[Zn^{2+}]$ and $[ZnY^{2-}]$,

$$\frac{k^+}{[ZnY^{2-}]_0} = \frac{k_1^+ + k_2^+[H^+] + k_3^+[H^+]\dfrac{1 + K_{ZnOAc}[OAc^-]}{[Zn^{2+}]_{0,app}}}{1 + \sum_{n=1}^{2} K_{Pb(OAc)_n}[OAc^-]^n} \quad (12)$$

respectively. Since both Zn^{2+} and ZnY^{2-} are present in the reaction mixture in a large excess over concentration of lead ions, the change of their concentrations during the reaction can be neglected.

The rate constants in Eq. 10 were calculated from the plots of $k^+/[ZnY^{2-}]_0$ against the hydrogen ion concentration, which are tabulated in Table I. The formation constants of acetatolead(II) and acetatozinc(II) complexes which were used in the calculation are those that were determined polarographically at ionic strength 0.2 at 15 and 25°C[5]. The concentration of uncomplexed acetate in the solution was calculated by successive approximation.

The rate constants of both elementary reactions 8b and 9b were calculated from the relations given in Eqs. 11, using the values of K^H_{ZnHY}, K_{ZnY} and K_{HY} which were obtained with the same procedure as described previously[3,4] and are given in Table II. The rate constants of three elementary reactions 7, 8b and 9b thus calculated are summarized in Table III.

TABLE I. RATE CONSTANTS IN EQ. 10 OBTAINED EXPERIMENTALLY AT IONIC STRENGTH 0.2 AT VARIOUS TEMPERATURES

Temp., °C	Rate constant, l. mol^{-1} sec^{-1}		
	k_1^+	k_2^+	k_3^+
15	4.6	6×10^5	1.1×10^4
25	8.9	9×10^5	2.4×10^4

TABLE II. EQUILIBRIUM CONSTANTS AT IONIC STRENGTH 0.2 AT VARIUS TEMPERATURES

Temp., °C	log K_{ZnY}	pK_{HY}	log K^H_{ZnHY}
15	16.17	10.20	—
20	—	—	2.9
25	15.94	10.04	—

TABLE III. RATE CONSTANTS FOR REACTIONS 7, 8b AND 9b AT IONIC STRENGTH 0.2 AT VARIOUS TEMPERATURES

Temp., °C	Rate constant, l. mol^{-1} sec^{-1}		
	k_1^+	k_{II}^+	k_{III}^+
15	4.6	8×10^2	1.0×10^{10}
25	8.9	1×10^3	1.9×10^{10}

From these results, the rate constants expressed in the form of the Arrhenius equation were derived and the heats of activation and entropies of activation were calculated for three elementary reactions 7, 8b and 9b. The results are given in Table IV.

Bril et al. determined the rate constants of this substitution reaction in acetate buffer solutions of ionic strength 1.0 containing 0.1 M acetate at 25°C[1], but they took no account of the effect of acetate ion on the reaction rate. Therefore, the rate constants reported by Bril et al., R_1, R_2 and R_3[6], correspond to those given by the equations,

$$R_1 = \frac{k_1^+}{1 + \sum_{n=1}^{2} K_{Pb(OAc)_n}[OAc^-]^n}$$

$$R_2 = \frac{k_2^+}{f_H\{1 + \sum_{n=1}^{2} K_{Pb(OAc)_n}[OAc^-]^n\}} \qquad (13)$$

$$R_3 = \frac{k_3^+\{1 + K_{ZnOAc}[OAc^-]\}}{f_H\{1 + \sum_{n=1}^{2} K_{Pb(OAc)_n}[OAc^-]^n\}}$$

where f_H means the activity coefficient of hydrogen ions. Since the values of K_{PbOAc}, $K_{Pb(OAc)_2}$, K_{ZnOAc}, K^H_{ZnHY}, K_{ZnY} and K_{HY} and f_H are all unknown at ionic strength 1.0 at 25°C, it seems impossible to calculate the rate constants R_1, R_2 and R_3 in Eq. 13, using the values of k_1^+, k_2^+ and k_3^+ obtained in this study.

In the measurement of the reaction rate, Bril et al. employed the streaming mercury electrode. They stated that the streaming electrode had an advantage over the conventional dropping mercury electrode when used as an indicator electrode for kinetic measurements. However, the diffusion current obtained with the streaming mercury electrode was considered less accurate and less reproducible than that obtained with the dropping mercury electrode. For comparison, therefore, the measurements were made under the same condition as Bril et al.'s but with the dropping mercury electrode, and the values of R_1, R_2 and R_3 were determined with the same procedure as mentioned above. The results are given in Table V and compared with the values reported by Bril et al. An agreement between both results seems to indicate no significant discrepancy between the two methods[7].

TABLE IV. RATE CONSTANTS, HEATS OF ACTIVATION AND ENTROPIES OF ACTIVATION FOR REACTIONS 7, 8b AND 9b

Reaction	Rate constant, expressed in l. mol^{-1} sec^{-1}	ΔH^*, kcal. mol^{-1}	ΔS^*, e. u.
$Pb^{2+} + ZnY^{2-} \rightarrow PbY^{2-} + Zn^{2+}$	$k_1^+ = 1.1\times10^9 \exp(-11{,}000/RT)$	10	−19
$Pb^{2+} + ZnHY^- \rightarrow PbHY^- + Zn^{2+}$	$k_{II}^+ = 9.3\times10^7 \exp(-6{,}700/RT)$	6.1	−24
$Pb^{2+} + HY^{3-} \rightarrow PbY^{2-} + H^+$	$k_{III}^+ = 1.2\times10^{10} \exp(-11{,}000/RT)$	10	22

6) R_1, R_2 and R_3 correspond to k_0^+, k_1^+ and k_2^+, respectively, given in the first paper of Bril, Bril and Krumholz[1].

7) The values of R_2 are considerably less accurate than those of R_1 and R_3[4].

TABLE V. RATE CONSTANTS EXPRESSED IN EQS. 13 OBTAINED EXPERIMENTALLY AT IONIC STRENGTH 1.0 AT 25°C

	Rate constant, l. mol^{-1} sec^{-1}, obtained	
	by Bril et al.	by the present authors
R_1	0.34	0.30
R_2	2×10^4	8×10^4
R_3	1.1×10^3	1.5×10^3

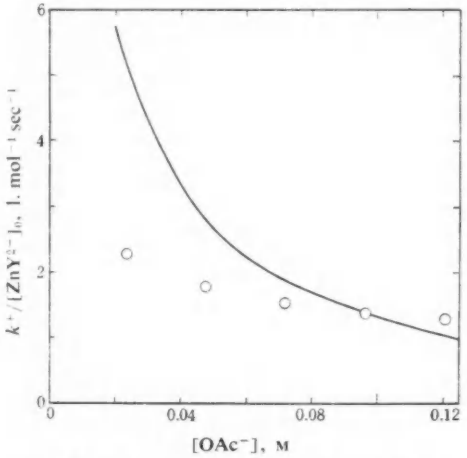


Fig. 2. $k^+/[ZnY^{2-}]_0$ as a function of the acetate concentration in acetate buffer–potassium nitrate solutions of pH 5.09 and ionic strength 0.2 at 15°C. Initial concentrations are: $[Pb^{2+}]_{0,app}=1.65\times10^{-4}$ M; $[ZnY^{2-}]_0=5.00\times10^{-3}$ M; $[Zn^{2+}]_{0,app}=1.00\times10^{-2}$ M. —— indicates the theoretical curve and ○, the experimental value.

From Eq. 10, it is predicted that the reaction rate decreases with increasing concentrations of acetate ions. Using Eq. 12, the values of $k^+/[ZnY^{2-}]_0$ were calculated at various acetate concentrations with the rate constants given in Table I and the formation constants of the acetato complexes. They are plotted against the acetate ion concentration as shown in Fig. 2. The overall rate constants ($k^+/[ZnY^{2-}]_0$) were actually measured at varied acetate concentrations under the conditions in which the theoretical values were calculated. They are also given in Fig. 2. Both the calculated and the experimental values increase with decreasing concentrations of aetate. The discrepancy existing between the calculated and experimental values is considered probably to be due to the contribution of the reaction in which acetato complexes participate.

Summary

The kinetic study of the substitution reaction of lead ions and zinc-EDTA complexs in acetate buffer solutions has been carried out with the same procedure as that reported previously[2-4]. Under the experimental conditions, the reaction is found to proceed simultaneously through three different reaction paths, and the rate constants for three elementary reactions of those reaction paths are determined at ionic strength 0.2 at 15 and 25°C. The heats of activation and the entropies of activation for the three elementary reactions are calculated.

The authors thank the Ministry of Education for the financial support granted for this research. Part of the experimental work has been carried out with the assistance of Dr. Gen Satô.

Department of Chemistry
Faculty of Science
Tohoku University
Katahira-cho, Sendai

Synthetic Studies on the Benzofuran Derivatives. VI. Synthesis of Furano(2'', 3'' : 7, 8)-2'-methoxy-isoflavone Derivatives

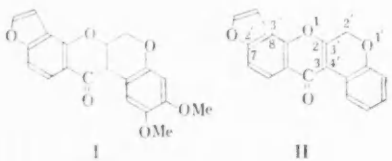
By Yoshiyuki KAWASE, Kiyomi OGAWA, Setsuko MIYOSHI and Kenji FUKUI*

(Received February 15, 1960)

The structure of elliptone, the simplest member of rotenoids, has been established[1] to be a dimethoxy derivative of furano-chromano-chromanone (I), but the synthesis of it and other rotenoids** is not yet achieved. Many efforts have hitherto been made on the synthesis of rotenoids, and recently the total synthesis of dihydrorotenone has been accomplished by Miyano and Matsui[2], who succeeded in the reduction of the chromenochromone system to give the chromanochromanone system. Therefore, it seems desirable for the synthesis of elliptone to prepare the furano-chromenochromone system, from which the furano-chromanochromanone system may be derived by reduction. For the preparation of chromenochromones, two methods are available —those are (a) by chromenochromone-ring formation of phenoxyacetic acid-2-(*o*-hydroxy)-acetophenones[3] and (b) by chromen-ring formation of 2-methyl[4] or 2-ethoxymethyl[5] derivatives of 2'-methoxy-isoflavones.

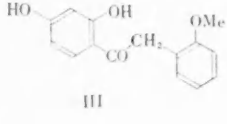
Following the latter method, the authors attempted to synthesize furano(2'', 3'' : 7, 8)-chromeno(3', 4' : 2, 3)chromone (II), and in this paper describe the preparation of three intermediate derivatives (VIIIa, b, c) of furano (2'', 3'' : 7, 8)-2'-methoxy-isoflavone.

Starting from three isoflavones (IVa, b, c), furan-ring was built up on the 7, 8-position of them smoothly by a similar method reported in the previous papers[6] as follows: the isoflavones IVa, b, c were formulated by the hexamine method to give 8-formyl derivatives (Va, b, c), which were converted into α-ethoxycarbonyl-furano derivatives (VIa, b, c) by the action of ethyl bromomalonate and potassium carbonate in acetone. Hydrolysis of the esters VIa, b, c under mild alkaline condition (aqueous sodium carbonate in acetone) furnished the corresponding acids VIIa, b, c which were decarboxylated in quinoline with the aid of copper catalyst

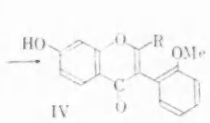


I

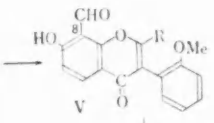
II



III

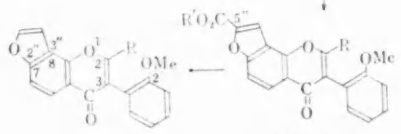


IV



V

For IV-VIII
a) R=-H
b) R=-Me
c) R=-CH_2OEt



VIII

VI R'=-Et
VII R'=-H

* Present adress: Faculty of Science, Hiroshima University, Hiroshima.

1) S. H. Harper, *J. Chem. Soc.*, **1939**, 1099, 1424; ibid., **1942**, 587.

** Recentry, total synthesis of *dl*-deguelin has been reported by H. Fukami, J. Oda, G. Sakata and M. Nakajima (*Bull. Agr. Chem. Soc. Japan*, **24**, 327 (1960)). (added in proof).

2) M. Miyano und M. Matsui, *Chem. Ber.*, **91**, 2044 (1958); ibid., **92**, 2487 (1959); M. Matsui und M. Miyano, *Proc. Japan Acad.*, **35**, 175 (1959).

3) A. Robertson, *J. Chem. Soc.*, **1933**, 489, 1163.

4) T. R. Seshadri and S. Varadarajan, *Proc. Indian Acad. Sci.*, **37A**, 784 (1953).

5) A. C. Mehta and T. R. Seshadri, ibid., **42A**, 192 (1955).

6) T. Matsumoto, Y. Kawase, M. Nanbu and K. Fukui, This Bulletin, **31**, 688 (1958); K. Fukui and Y. Kawase, ibid., **31**, 693 (1958).

to give furano(2'', 3'' : 7, 8)-2'-methoxy-isoflavone*** (VIIIa), furano(2'', 3'' : 7, 8)-2'-methoxy-2-methylisoflavone (VIIIb) and furano-(2'', 3'' : 7, 8)-2'-methoxy-2-ethoxymethylisoflavone (VIIIc) respectively. Attempts to prepare II from VIIIb or VIIIc are now in progress.

Experimental[7]

2, 4-Dihydroxyphenyl 2-Methoxybenzyl Ketone (III).—This ketone was prepared from resorcinol and *o*-methoxyphenylacetonitrile by the Hoesch reaction[4] or more conveniently from resorcinol and *o*-methoxy-phenylacetic acid by means of boron trifluoride[5]. M. p. 158~159°C (reported m. p. 159~160°C[4]).

7-Hydroxy-2'-methoxy-isoflavone (IVa).—A mixture of the ketone III (10 g.), ethyl orthoformate (20 ml.), piperidine (4 ml.) and pyridine (80 ml.) was refluxed for 8 hr. in an oil-bath. The resulting solution was cooled and acidified with dilute hydrochloric acid. The crystals formed were recrystallized from ethanol to give IVa, m. p. 226.5~227.5°C (reported m. p. 230~232°C[8]), having negative ferric reaction in ethanol; yield 6.6 g.

7-Hydroxy-2'-methoxy-8-formyl-isoflavone (Va).—A mixture of the isoflavone IVa (6 g.), hexamine (32 g.) and acetic acid (100 ml.) was heated in an oil bath (100°C) for 6 hr. The resulting hot solution was treated with hot dilute hydrochloric acid (1 : 1, 100 ml.) and again heated for 10 min., then allowed to stand overnight. The precipitates were collected, washed with water and dissolved in benzene, and the benzene solution was extracted with dilute aqueous sodium hydroxide. The product obtained by acidifying the alkaline solution was recrystallized from ethanol to give Va in faintly yellow colored microcrystals, m. p. 187~189°C, having a red ferric reaction in ethanol; yield 2.1 g.

Found: C, 68.40; H, 4.25. Calcd. for $C_{17}H_{12}O_5$: C, 68.92; H, 4.05%.

7-Hydroxy-2'-methoxy-8-formyl-2-methylisoflavone (Vb).—By the similar method as described for Va, 1.3 g. of Vb was obtained from 7-hydroxy-2'-methoxy-2-methylisoflavone[4] (IVb, m. p. 222~223°C, 3 g.), hexamine (16 g.) and acetic acid (50 ml.); m. p. 165.5~166.5°C (from ethanol) faintly yellow colored microcrystals, having a red ferric reaction in ethanol.

Found: C, 69.18; H, 4.67. Calcd. for $C_{18}H_{14}O_5$: C, 69.68; H, 4.52%.

7-Hydroxy-2'-methoxy-8-formyl-2-ethoxymethyl-isoflavone (Vc).—By the similar method, 1 g. of Vc was obtained from 7-hydroxy-2'-methoxy-2-ethoxymethylisoflavone[5] (IVc, m. p. 163~164°C, 1.5 g.), hexamine (8 g.) and acetic acid (25 ml.); m. p. 129~130°C (from 80% ethanol) faintly yellow colored microcrystals, having a red ferric reaction in ethanol.

Found: C, 68.04; H, 5.24. Calcd. for $C_{20}H_{18}O_6$: C, 67.79; H, 5.12%.

5''-Ethoxycarbonyl-furano(2'',3'':7,8)-2'-methoxy-isoflavone (VIa).—To a solution of the formyl-isoflavone Va (2 g.) and ethyl bromomalonate (2.2 g.) in anhydrous acetone (65 ml.) was added anhydrous potassium carbonate (7.8 g.), and the mixture was refluxed for 8 hr. in a steam bath. The resulting solution was filtered from precipitates and the solvent was distilled off. The residual oily product was recrystallized from ethanol to give VIa, m. p. 164.5°C, having a negative ferric reaction; yield 1.1 g.

Found: C, 69.12; H, 4.50. Calcd. for $C_{21}H_{16}O_6$: C, 69.23; H, 4.40%.

5''-Ethoxycarbonyl-furano(2'',3'':7,8)-2'-methoxy-2-methylisoflavone (VIb).—By the similar method as described for VIa, 1.2 g. of VIb was obtained from the formyl-isoflavone Vb (2 g.), ethyl bromomalonate (2.2 g.) and potassium carbonate (7.8 g.); m. p. 161.5~162.5°C (from ethanol), having a negative ferric reaction.

Found: C, 69.97; H, 4.96. Calcd. for $C_{22}H_{18}O_6$: C, 69.84; H, 4.76%.

5''-Ethoxycarbonyl-furano(2'',3'':7,8)-2'-methoxy-2-ethoxymethylisoflavone (VIc).—By the similar method, 1 g. of VIc was obtained from the formyl-isoflavone Vc (1.5 g.), ethyl bromomalonate (1.7 g.) and potassium carbonate (6 g.); m. p. 177.5~178.5°C (from ethanol), having a negative ferric reaction.

Found: C, 68.55; H, 5.33. Calcd. for $C_{24}H_{22}O_7$: C, 68.24; H, 5.25%.

5''-Carboxy-furano(2'', 3'' : 7, 8)-2'-methoxy-isoflavone (VIIa).—A mixture of the ester VIa (0.8 g.), acetone (40 ml.) and 5% aqueous sodium carbonate (30 ml.) was refluxed for 3 hr. Dilute hydrochloric acid (20 ml.) was added to the resulting mixture, and acetone was distilled off. The precipitates were collected and recrystallized from ethanol; m. p. 273.5~274°C, having a negative ferric reaction; yield 0.6 g.

Found: C, 67.83; H, 3.65. Calcd. for $C_{19}H_{12}O_6$: C, 67.86; H, 3.57%.

5''-Carboxy-furano(2'', 3'' : 7, 8)-2'-methoxy-2-methylisoflavone (VIIb).—By the similar method as described for VIIa, 1.1 g. of VIIb was obtained from the ester VIb (1.2 g.), acetone (50 ml.) and 5% aqueous sodium carbonate (50 ml.); m. p. 294°C (from ethanol), having a negative ferric reaction.

Found: C, 68.41; H, 4.02. Calcd. for $C_{20}H_{14}O_6$: C, 68.57; H, 4.00%.

5''-Carboxy-furano(2'', 3'' : 7, 8)-2'-methoxy-2-ethoxymethylisoflavone (VIIc).—By the similar method, 0.3 g. of VIIc was obtained from the ester VIc (0.5 g.), acetone (20 ml.) and 5% aqueous sodium carbonate (20 ml.); m. p. 233~234°C (from ethanol), having a negative ferric reaction.

Found: C, 67.21; H, 4.57. Calcd. for $C_{22}H_{18}O_7$: C, 67.00; H, 4.60%.

Furano(2'',3'' : 7,8)-2'-methoxy-isoflavone (VIIIa).—A mixture of the acid VIIa (0.5 g.), copper powder (0.3 g.) and quinoline (20 ml.) was heated while stirring in an atmosphere of nitrogen for 20 min.

*** It has recently been reported that munetone is 5''-isopropyl derivative of VIIIa by N. L. Dutta (*J. Indian Chem. Soc.*, **36**, 165 (1959)).

7) Melting points are uncorrected, and the products are colorless microcrystals unless otherwise noted.

8) Private communication from H. Fukami of the Agricultural Department of Kyoto University.

9) P. K. Grover and T. R. Seshadri, *Proc. Indian Acad. Sci.*, **38A**, 122 (1953).

at 190~200°C until evolution of carbon dioxide ceased. Quinoline was distilled off by steam-distillation, and the residual product was dissolved in ethyl acetate, filtered from copper, and washed with dilute sulfuric acid and dilute aqueous sodium hydroxide. Evaporation of ethyl acetate gave VIIIa, m. p. 182°C (from ethyl acetate and then from ethanol), having a negative ferric reaction; yield 0.3 g.

Found: C, 74.04; H, 4.20. Calcd. for $C_{18}H_{12}O_4$: C, 73.97; H, 4.11%.

Furano(2'',3'':7,8)-2'-methoxy-2-methylisoflavone (VIIIb).—By the similar method as described for VIIIa, 0.6 g. of VIIIb was obtained from the acid VIIb (1 g.), copper powder (0.5 g.) and quinoline (35 ml.); m. p. 168°C (from ethyl acetate and then from ethanol), having a negative ferric reaction.

Found: C, 74.33; H, 4.64. Calcd. for $C_{19}H_{14}O_4$: C, 74.51; H, 4.58%.

Furano(2'', 3'' : 7, 8)-2'-methoxy-2-ethoxymethylisoflavone (VIIIc).—By the similar method, 0.5 g. of VIIIc was obtained from the acid VIIc (1 g.), copper powder (0.5 g.) and quinoline (35 ml.); m. p. 119~120°C (from ethyl acetate and then from ethanol), having a negative ferric reaction.

Found: C, 72.18; H, 5.21. Calcd. for $C_{21}H_{18}O_5$: C, 71.99; H, 5.18%.

The authors are grateful to the Faculty of Pharmacy of this University for microanalyses, and the cost of this research was partly defrayed from the Research Grant of the Ministry of Education, to which the author's thanks are also due.

Faculty of Literature and Science
Toyama University
Hasumachi, Toyama

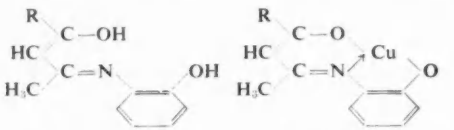
Metal Complexes with Terdentate Ligands. II[1]. *Synthesis and Properties of Tri-coordinated Copper(II) Complexes*[2]

By Yoneichiro MUTO

(Received February 13, 1960)

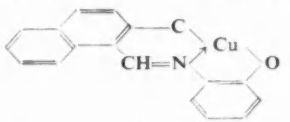
So-called tri-coordinated copper(II) complexes have been reported by several investigators. For instance, Pfeiffer et al.[3] have prepared salicylaldehyde-2-hydroxyanil copper(II), while Yamaguchi[4], Dikmen et al.[5] and the present author[6] have synthesized salicylaldehyde-2-hydroxybenzylimine copper(II). Although it is presumed that the copper atoms in these chelates have an unusual coordination number, three, it is desirable to study extensively tri-coordinated copper(II) complexes having different types of ligand molecules in order to obtain more definite information about the tri-coordination of the copper atoms. This paper presents the synthesis and properties of some tri-coordinated copper(II) complexes.

I. Copper(II) Complexes having Schiff Bases of 2-Aminophenol.—In order to obtain terdentate ligands of similar type to salicylaldehyde-2-hydroxyanil, the present author prepared acetylacetone-mono-(2-hydroxyanil) and benzoylacetone-mono-(2-hydroxyanil) by the condensation of 2-aminophenol with acetylacetone and benzoylacetone, respectively. It is expected that they act as terdentate ligands in their enol forms shown in A′ and B′. That is to say, they have two hydrogen atoms capable of being replaced by metal atoms (i.e., hydrogen



A′ R: CH_3
B′ R: C_6H_5*

A R: CH_3
B R: C_6H_5



C

1) Y. Muto, This Bulletin, **31**, 1017 (1958).
2) This was presented partly in a previous paper (Y. Muto, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **76**, 1405 (1955).), and partly at the Symposium on Co-ordination Compounds held by the Chemical Society of Japan in Fukuoka on October 21, 1959.
3) P. Pfeiffer, Th. Hesse, G. Pfizner, W. Scholl and H. Thielert, *J. prakt. Chem.*, **149**, 217 (1937).
4) M. Yamaguchi, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **74**, 261 (1954).
5) C. Dikmen and T. Günduz, *Ber.*, **89**, 2637 (1956).
6) Y. Muto, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **74**, 274 (1953).

* It is known that primary amines condense with benzoylacetone to form Schiff bases having a general formula $C_6H_5CO{\cdot}CH_2C(=N-R){\cdot}CH_3$. Therefore, benzoylacetone-mono-(2-hydroxyanil) must have the structure shown in B′, althogh it may be presumed that the anil has the structure, $C_6H_5C(=N-C_6H_4OH){\cdot}CH_2CO{\cdot}CH_3$.

of the phenolic and enolic hydroxyl groups) as well as a nitrogen atom capable of being coordinated to a metal atom (i.e., nitrogen of an azomethine group). Therefore, they will react with copper(II) ion to form chelates having a five-membered ring and a six-membered one, which are very common. On being treated with copper(II) acetate, they yielded complexes having the compositions expected for the formulae A and B. Similarly, 2-hydroxynaphthaldehyde-(1)-[*o*-hydroxyanil] yielded a copper(II) complex having the composition shown in C.

These complexes were readily soluble in pyridine, slightly soluble in nitrobenzene, and soluble with difficulty in other common organic solvents. When these complexes were recrystallized from pyridine, the corresponding monopyridinates were formed. The pyridinates liberated pyridine gradually at 70°C and completely at 120°C. The decrease in weight corresponded to one molecule of pyridine per one atom of copper. From these facts, it seems very likely that these pyridine-free complexes have two interlocked chelate rings shown in A, B and C, although their structures might not be so simple as is implied by the ordinary chemical formulae. This conclusion has been suggested further by magnetochemical studies[7,8]. As the present author et al. have already reported in previous papers[7,8], these pyridine-free complexes show magnetic moments considerably smaller than the theoretical value predicted from the presence of one odd electron, and the subnormal moments suggest the presence of dimeric molecules, in which two copper atoms are very close to each other as in copper(II) acetate monohydrate crystals, whereas the pyridinates have the normal moments expected for tetra-coordinated copper(II) complexes (Table I).

It has been presumed by the method of ligand-exchange[9] that these tri-coordinated complexes are more stable than the tetra-coordinated complexes having ligands similar to A′ or B′ from the structural point of view, but that they are less stable than bis-8-oxyquinolin copper(II) or salicylaldehyde-ethylenediimine copper(II) (Table II). For instance, when an alcoholic solution of benzoylacetone-mono-(2-hydroxyanil) was mixed with an alcoholic solution of bis-salicylaldehydeanil copper(II), benzoylacetone-mono-(2-hydroxyanil) copper(II) was formed, whereas the reverse reaction of this exchange did not take place.

II. Copper(II) Complexes Having Schiff Bases of Anthranilic Acid. — These complexes were synthesized by essentially the same method as that described in I. Namely, anthranilic acid was subjected to condensation with 5-*X*-salicylaldehyde (*X*-: chloro-, bromo- or nitro-), and the resulting Schiff bases (Formula D′) were treated with copper(II) acetate. In this

Table I. The magnetic moments of tri-coordinated copper(II) complexes and their pyridinates

Chelates	μ (B. M.)	Reference
Salicylaldehyde-2-hydroxybenzylimine Cu(II)	0.87	7
Monopyridino-salicylaldehyde-2-hydroxybenzylimine Cu(II)	1.83	7
Salicylaldehyde-2-hydroxyanil Cu(II)	1.34	7
Monopyridino-salicylaldehyde-2-hydroxyanil Cu(II)	1.84	7
Acetylacetone-mono-(2-hydroxyanil) Cu(II)	1.37	7
Monopyridino-acetylacetone-mono-(2-hydroxyanil) Cu(II)	1.81	7
Benzoylacetone-mono-(2-hydroxyanil) Cu(II)	1.11	7
Monopyridino-benzoylacetone-mono-(2-hydroxyanil) Cu(II)	1.84	7
2-Hydroxynaphthaldehyde-(1)-[*o*-hydroxyanil] Cu(II)	1.39	8
Monopyridino-2-hydroxynaphthaldehyde-(1)-[*o*-hydroxyanil] Cu(II)	1.77	8
5-Bromosalicylalanthranilic acid Cu(II)	1.87	8
Monopyridino-5-bromosalicylalanthranilic acid Cu(II)	1.86	8
5-Nitrosalicylalanthranilic acid Cu(II)	1.87	8
Monopyridino-5-nitrosalicylalanthranilic acid Cu(II)	1.84	8
Salicylaldehyde-2-mercaptoanil Cu(II)	1.82*	
2-Hydroxy-5-methylbenzaldehyde-*o*-mercaptoanil Cu(II)	1.93*	
5-Bromosalicylaldehyde-*o*-mercaptoanil Cu(II)	1.76*	

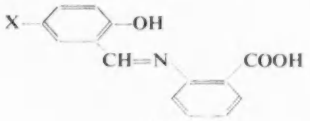
* The magnetic susceptibilities were determined by the Gouy method at room temperature (28~31°C).

7) M. Kishita, Y. Muto and M. Kubo, *Naturwissenschaften*, **44**, 372 (1957); *Australian J. Chem.*, **10**, 386 (1957).
8) M. Kishita, Y. Muto and M. Kubo, *Naturwissenschaften*, **44**, 612 (1957); *Australian J. Chem.*, **11**, 309 (1958).
9) Y. Muto, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **76**, 252 (1955).

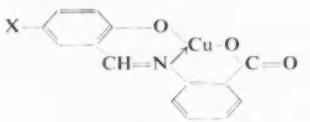
TABLE II. THE LIGAND-EXCHANGE REACTIONS

Chelates	Replacing agents	Resulting chelates	Analysis Cu % Calcd.	Cu % Found	N % Calcd.	N % Found
Bis-acetylacetone Cu(II)	Acetylacetone-mono-(2-hydroxyanil)	Acetylacetone-mono-(2-hydroxyanil) Cu(II)	25.14	25.08	5.54	5.71
Bis-benzoylacetone Cu(II)	Benzoylacetone-mono-(2-hydroxyanil)	Benzoylacetone-mono-(2-hydroxyanil) Cu(II)	20.18	20.15	4.45	4.49
Bis-2-hydroxynaphthaldehyde-(1) Cu(II)	2-Hydroxynaphthaldehyde-(1)-[*o*-hydroxyanil]	2-Hydroxynaphthaldehyde-(1)-[*o*-hydroxyanil] Cu(II)	19.52	19.38	4.40	4.54
Bis-salicylaldimine Cu(II)	Benzoylacetone-mono-(2-hydroxyanil)	Benzoylacetone-mono-(2-hydroxyanil) Cu(II)	20.18	20.09	4.45	4.54
Bis-salicylaldehyde-anil Cu(II)	Acetylacetone-mono-(2-hydroxyanil)	Acetylacetone-mono-(2-hydroxyanil) Cu(II)	25.14	25.16	5.54	5.64
Bis-dimethylglyoxime Cu(II)	Benzoylacetone-mono-(2-hydroxyanil)	Benzoylacetone-mono-(2-hydroxyanil) Cu(II)	20.18	20.23	4.45	4.57
Bis-ethylacetoacetate Cu(II)	Acetylacetone-mono-(2-hydroxyanil)	Acetylacetone-mono-(2-hydroxyanil) Cu(II)	25.14	25.13	5.54	5.67
Benzoylacetone-mono-(2-hydroxyanil) Cu(II)	8-Oxyquinoline	Bis-8-oxyquinoline Cu(II)	18.06	18.11	7.96	8.02
Acetylacetone-mono-(2-hydroxyanil) Cu(II)	Salicylaldehyde-ethylene-diimine	Salicylaldehyde-ethylene-diimine Cu(II)	19.27	19.20	8.49	8.61

way the complexes having the compositions shown in D were obtained. These complexes were soluble with difficulty in common organic solvents such as ethanol or benzene, while their monopyridinates were soluble in ethanol to a certain extent. Pyridine could be removed from the pyridinates by heating at about 150～200°C in air or boiling in ethanol.



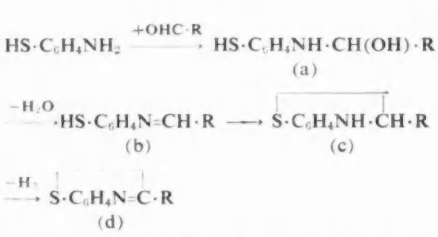
D'



D

X: Cl, Br, NO_2

III. Copper(II) Complexes Having Schiff Bases of 2-Aminothiophenol.—Unlike the case of 2-aminophenol or anthranilic acid, 2-aminothiophenol condenses with aldehydes to form benzothiazoles. It seems probable from the results of several investigators[10] that the mechanism of the reaction can be represented as follows:



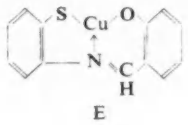
Although the thiazolines (c) were isolated as intermediates in the formation of the thiazoles (d) from 2-aminothiophenol and aldehydes, neither the aldols (a) nor the anils (b) have been isolated.

Claasz[10] obtained 2-(*o*-hydroxyphenyl)benzothiazol (m. p. 130°C) by the reaction of 2-aminothiophenol with salicylaldehyde. Recently, Charles and Freiser[11] have obtained 2-(*o*-hydroxyphenyl)benzothiazoline (m.p. 136～137°C). However, in the course of this study, the present author has mixed 2-aminothiophenol and salicylaldehyde in ethanol and obtained a product of m. p. 148～148.5°C. This compound was converted into 2-(*o*-hydroxyphenyl)benzothiazol when it was recrystallized from ethanol containing a small amount of hydrogen chloride. On oxidation with hydrogen peroxide, this compound yielded 2-salicylaminophenyldisulfide, $(HO\cdot C_6H_4CH{=}N\cdot C_6H_4S{-})_2$, indicating the presence of a mercapto group. The elementary analysis of this compound agreed with the formula $HO\cdot C_6H_4CH{=}N\cdot C_6H_4SH$. On treating this compound with

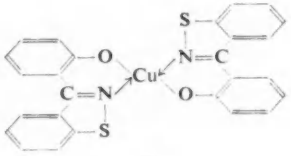
10) M. Claasz, *Ber.*, **45**, 1031 (1912); **49**, 1141 (1916); M. T. Bogert and A. Stull, *J. Am. Chem. Soc.*, **47**, 3038 (1925); H. P. Lankelma and P. X. Sharnoff, ibid., **53**, 2654 (1931); M. T. Bogert and B. Naiman, ibid., **57**, 1529 (1935).

11) R. G. Charles and H. Freiser, *J. Org. Chem.*, **18**, 422 (1953).

copper(II) acetate, it yielded a complex having a composition agreeing with the formula E.



E



F

This complex was also prepared by the condensation of bis-salicylaldehyde copper(II) with 2-aminothiophenol. This formula was suggested further by the following observations. First, when this complex was decomposed by dilute sulfuric acid, salicylaldehyde was liberated. If this complex had had a thiazol or thiazoline link, salicylaldehyde would not have been liberated from it. Secondly, when this complex was recrystallized from pyridine, its monopyridinate was formed in the same manner as the tri-coordinated copper(II) complexes described in I and II. This shows that the composition of the complex has the molar ratio 1 : 1 of copper(II) to the ligand molecule.

As an example of metal complexes having the Schiff base of 2-aminothiophenol, only the following zinc complex has been reported. Bogert and Naiman[10] have shown that zinc 2-aminothiophenolate could be condensed directly with aromatic aldehydes to form zinc complexes having a general formula $(R\cdot CH{=}N\cdot C_6H_4S{-})_2Zn$, which yielded 2-R-benzothiazoles when they were boiled in glacial acetic acid. The present author has also found that the corresponding 2-R-benzothiazoline does not form zinc or copper(II) complexes, while 2-(*o*-hydroxyphenyl)benzothiazol forms a copper(II) complex having the molar ratio 1 : 2 of copper(II) to the ligand molecule as shown in F. This result suggests that the thiazolines (c) cannot be converted into anils (b) by the action of metal ions. Therefore, the fact that reaction product of 2-aminothiophenol with salicylaldehyde (m. p. 148~148.5°C) yielded a complex having the composition shown in E gives support to the conclusion that the product must be salicylaldehyde-2-mercaptoanil.

2-Hydroxy-5-methylbenzaldehyde-*o*-mercaptoanil copper(II), 5-bromosalicylaldehyde-*o*-mercaptoanil copper(II) and their monopyridinates could be prepared in an analogous way.

12) W. Suida, *J. prakt. Chem.*, [2] **83**, 241 (1911).

Experimental

Preparation of Copper(II) Complexes.—Appropriate amines were subjected to condensation with suitable ketones or aldehydes in ethanol. The resulting Schiff bases were separated and were treated with copper(II) acetate in the mixture of ethanol and water. The crude products which precipitated were separated, dissolved in pyridine and filtered. When water was added to the filtrates, the monopyridinates of the complexes were formed as crystals. On heating to an appropriate temperature (100~200°C), almost all the pyridine was removed. To ensure the complete removal, the residue were boiled in ethanol on a water bath for about half an hour. After being allowed to cool, the precipitates were filtered and dried at about 90°C. In this way pyridine-free copper(II) complexes were prepared. None of them melted below 300°C.

(1) **Acetylacetone-mono-(2-hydroxyanil) Copper(II).**—This was prepared from copper(II) acetate and acetylacetone-mono-(2-hydroxyanil)[12] (m. p. 187°C, yellow prismatic crystals. Found: N, 7.48. Calcd. for $C_{11}H_{13}O_2N$: N, 7.33%). The copper(II) chelate formed green prismatic crystals, while its pyridinate crystallized as deep green plates.

Found: N, 5.78; Cu, 25.32. Calcd. for $C_{11}H_{11}\cdot O_2NCu$: N, 5.54; Cu, 25.14%. Found: N, 8.69; Cu, 19.34; py., 23.73. Calcd. for $C_{11}H_{11}O_2NCu\cdot C_5H_5N$, N, 8.44; Cu, 19.15; py., 23.84%.

(2) **Benzoylacetone-mono-(2-hydroxyanil) Copper(II).**—Benzoylacetone-mono-(2-hydroxyanil) was prepared by the condensation of benzoylacetone (1.7 g.) with 2-aminophenol (1.1 g.) in ethanol (30 ml.). It was isolated in the form of yellow prismatic crystals; m. p. 169~170 C, yield 2 g.

Found: C, 75.76; H, 6.02; N, 5.81. Calcd. for $C_{16}H_{13}O_2N$: C, 75.87; H, 5.97; N, 5.53%.

The copper(II) chelate formed yellow-green needles, while its pyridinate crystallized as deep green plates.

Found: N, 4.49; Cu, 20.19. Calcd. for $C_{16}H_{11}\cdot O_2NCu$: N, 4.45; Cu, 20.18%. Found: N, 7.01; Cu, 16.23; py., 19.69. Calcd. for $C_{16}H_{13}O_2NCu\cdot C_5H_5N$: N, 7.09; Cu, 16.09; py., 20.03%.

(3) **2-Hydroxynaphthaldehyde-(1)-[*o*-hydroxyanil] Copper(II).**—This was prepared from copper(II) acetate and 2-hydroxynaphthaldehyde-(1)-[*o*-hydroxyanil][13] (m.p. 248~249°C, brown-red needles, Found: N, 5.41. Calcd. for $C_{17}H_{13}O_2N$: N, 5.32%). The copper(II) chelate was yellow-green, while its pyridinate was brown. Both crystallized as needles.

Found: N, 4.40; Cu, 19.52. Calcd. for $C_{17}H_{11}\cdot O_2NCu$: N, 4.31; Cu, 19.56%. Found: N, 6.99; Cu, 15.71; py., 19.39. Calcd. for $C_{17}H_{11}O_2NCu\cdot C_5H_5N$: N, 6.94; Cu, 15.73; py., 19.59%.

(4) **5-Chlorosalicylalanthranilic Acid Copper(II).**—5-Chlorosalicylalanthranilic acid was prepared by the condensation of 5-chlorosalicylaldehyde (1.6 g.) with anthranilic acid (1.4 g.) in ethanol (50 ml.). It was isolated in the form of orange prismatic crystals; m. p. 200°C, yield 2 g.

13) A. Senier and R. Clarke, *J. Chem. Soc.*, **99**, 2082 (1911—2).

Found: C, 61.10; H, 3.63; N, 5.39. Calcd. for $C_{14}H_{10}O_3NCl$: C, 60.99; H, 3.66; N, 5.08%.

The copper(II) chelate formed greenish brown needles, while its pyridinate crystallized as bright green plates.

Found: N, 4.29; Cu, 18.63. Calcd. for $C_{14}H_8O_3NClCu$: N, 4.15; Cu, 18.84%. Found: N, 6.83; Cu, 14.98; py., 19.24. Calcd. for $C_{14}H_8O_3NClCu \cdot C_5H_5N$: N, 6.73; Cu, 15.26; py., 19.00%.

(5) 5-Bromosalicylalanthranilic Acid Copper(II).—5-Bromosalicylalanthranilic acid was prepared by the condensation of 5-bromosalicylaldehyde (2 g.) with anthranilic acid (1.5 g.) in ethanol (50 ml.). It was isolated in the form of orange prismatic crystals; m. p. 202°C, yield 3 g.

Found: C, 52.47; H, 3.12; N, 4.48. Calcd. for $C_{14}H_{10}O_3NBr$: C, 52.52; H, 3.15; N, 4.38%.

The copper(II) chelate formed greenish brown needles, while its pyridinate crystallized as bright green plates.

Found: N, 3.96; Cu, 16.87. Calcd. for $C_{14}H_8O_3NBrCu$: N, 3.71; Cu, 16.82%. Found: N, 6.30; Cu, 13.86; py., 17.36. Calcd. for $C_{14}H_8O_3NBrCu \cdot C_5H_5N$: N, 6.08; Cu, 13.79; py., 17.17%.

(6) 5-Nitrosalicylalanthranilic Acid Copper(II).—5-Nitrosalicylalanthranilic acid was prepared by the condensation of 5-nitrosalicylaldehyde (1.7 g.) with anthranilic acid (1.4 g.) in ethanol (50 ml.). It was isolated in the form of red prismatic crystals; m. p. 271°C, yield 2.2 g.

Found: C, 58.67; H, 3.48; N, 10.04. Calcd. for $C_{14}H_{10}O_5N_2$: C, 58.74; H, 3.52; N, 9.79%.

The copper(II) chelate was greenish brown, while its pyridinate was bright green. Both formed prismatic crystals.

Found: N, 7.80; Cu, 18.43. Calcd. for $C_{14}H_8O_5N_2Cu$: N, 8.06; Cu, 18.27%. Found: N, 10.12; Cu, 15.15; py., 18.31. Calcd. for $C_{14}H_8O_5N_2Cu \cdot C_5H_5N$: N, 9.85; Cu, 14.89; py., 18.53%.

(7) Salicylaldehyde-2-mercaptoanil Copper(II).

—*a*) *Salicylaldehyde-2-mercaptoanil.*—Equimolar amounts (0.01 mol.) of 2-aminothiophenol and salicylaldehyde were dissolved in ethanol (30 ml.). The resulting solution was kept below 30°C by means of a cold water bath for about two hours; meanwhile the crude product precipitated. Recrystallization from ether gave colorless prismatic crystals; m. p. 148~148.5°C, yield 1.5 g.

Found: C, 68.27; H, 5.11; N, 6.08. Calcd. for $C_{13}H_{11}NOS$: C, 68.10; H, 4.84; N, 6.11%.

b) The copper(II) chelate was prepared by the reaction of salicylaldehyde-2-mercaptoanil described above with copper(II) acetate. This chelate was also obtained by the condensation of bis-salicylaldehyde copper(II) with 2-aminothiophenol in the same way as that described in (9). The pyridine-free complex was deep green, while the pyridinate was brown. Both crystallized as needles.

Found: C, 53.74; H, 3.41; N, 4.67; Cu, 21.80. Calcd. for $C_{13}H_9NOSCu$: C, 53.69; H, 3.12; N, 4.82; Cu, 21.85%. Found: N, 7.70; Cu, 17.31; py., 21.39. Calcd. for $C_{13}H_9NOSCu \cdot C_5H_5N$: N, 7.57; Cu, 17.18; py., 21.38%.

c) *Oxidation of Salicylaldehyde-2-mercaptoanil.*—Salicylaldehyde-2-mercaptoanil (0.5 g.) was dissolved in a mixture of ethanol (20 ml.) and aqueous hydrogen peroxide (30%, 5 ml.). After being heated gently on a water bath for ten minutes, the solution was allowed to stand overnight in air at room temperature; meanwhile yellow crystals precipitated. Recrystallization from benzene gave yellow needles; m. p. 171°C, yield 0.3 g. When this product was mixed with 2-salicylaminophenyldisulfide, the melting point was unaltered.

Found: N, 6.26. Calcd. for $C_{26}H_{16}O_2N_2S_2$: N, 6.19%.

(8) 5-Bromosalicylaldehyde-*o*-mercaptoanil Copper(II).—5-Bromosalicylaldehyde-*o*-mercaptoanil was prepared by the condensation of 5-bromosalicylaldehyde (2 g.) with 2-aminothiophenol (1.3 g.) in the same way as that described in (7). It was isolated in the form of colorless prismatic crystals; m. p. 147°C, yield 2 g.

Found: C, 50.90; H, 3.60; N, 4.31. Calcd. for $C_{13}H_{10}NOSBr$: C, 50.66; H, 3.27; N, 4.54%.

The copper(II) chelate was deep green, while its pyridinate was brown-red. Both crystallized as needles.

Found: C, 41.89; H, 2.42; N, 3.84; Cu, 17.25. Calcd. for $C_{13}H_8NOSBrCu$: C, 42.23; H, 2.18; N, 3.79; Cu, 17.19%. Found: N, 6.42; Cu, 14.30; py., 17.66. Calcd. for $C_{13}H_8NOSBrCu \cdot C_5H_5N$: N, 6.24; Cu, 14.16; py., 17.62%.

(9) 2-Hydroxy-5-methylbenzaldehyde-*o*-mercaptoanil Copper(II).—A solution of 2-aminothiophenol (0.7 g.) in ethanol (10 ml.) was added to a boiling solution of bis-2-hydroxy-5-methylbenzaldehyde copper(II) (1.7 g.) in ethanol (50 ml.). Deep green micro-crystals immediately precipitated. It was purified by recrystallization from pyridine. The copper(II) chelate was deep green, while its pyridinate was brown-red. Both crystallized as needles.

Found: C, 55.10; H, 3.87; N, 4.47; Cu, 20.85. Calcd. for $C_{14}H_{11}NOSCu$: C, 55.16; H, 3.64; N, 4.59; Cu, 20.84%. Found: N, 7.45; Cu, 16.59; py., 20.63. Calcd. for $C_{14}H_{11}NOSCu \cdot C_5H_5N$: N, 7.30; Cu, 16.55; py., 20.60%.

(10) Bis-2-(*o*-hydroxyphenyl)benzothiazol Copper(II).—A solution of copper(II) acetate monohydrate (0.5 g.) in water (30 ml.) was added to a boiling solution of 2-(*o*-hydroxyphenyl)benzothiazol (0.6 g.) in ethanol (30 ml.). Dark green microcrystals immediately precipitated. Recrystallization from pyridine gave dark green plates. Yield 0.7 g.

Found: N, 5.40; Cu, 12.38. Calcd. for $C_{26}H_{16}O_2N_2S_2Cu$: N, 5.31; Cu, 12.31%.

Bis-2-(*o*-hydroxy-*m*-bromophenyl)benzothiazol copper(II) could be prepared in a similar manner. It was isolated in the form of dark green plates.

Found: N, 4.21; Cu, 9.52. Calcd. for $C_{26}H_{14}O_2N_2S_2Br_2Cu$: N, 4.16; Cu, 9.43%.

Decomposition of Complexes (7), (8) and (9).—The complexes described in (7), (8) and (9) were decomposed by dilute sulfuric acid and then subjected to steam distillation. Salicylaldehyde, 5-bromosalicylaldehyde and 2-hydroxy-5-methylbenzaldehyde in the distillates were identified by their phenylhydrazones which melted at 143°C, 151°C and 149°C respectively.

Reaction of 2-R-benzothiazolines with Zinc or Copper(II) Acetate.—The reaction of 2-phenylbenzothiazoline with zinc or copper(II) acetate was

attempted in organic solvents such as ethanol, pyridine or acetic acid, but the product isolated was only 2-phenylbenzothiazol. Similarly, 5-chloro-2-phenylbenzothiazoline did not form zinc or copper-(II) complex.

Exchange Experiments.—Calculated amounts of appropriate reagents and suitable complexes were dissolved or suspended in ethanol. The mixtures were heated on a water bath for about five minutes. The resulting complexes were separated, washed with hot ethanol repeatedly and then dried at 80°C. The elementary analysis of the resulting complexes is presented in Table II.

The author is grateful to Mr. Michihiko Kishita of Nagoya University for measurements of magnetic susceptibilities and to Mr. Michio Shido of Kyushu University for elementary analyses.

Chemical Laboratory, Faculty of Liberal Arts, Saga University
Honjo-machi, Saga

Tropones. V. Addition Reaction of Tropone and Maleic Anhydride***

By Tetsuo NOZOE, Toshio MUKAI, Tsuneyuki NAGASE and Yasuo TOYOOKA

(Received March 15, 1960)

Troponoid compounds were at first considered to be resistant to the Diels-Alder type addition reactions because of their high aromaticity[1,2] but later, Sebe and others[3] found that hinokitiol underwent an addition reaction with maleic anhydride. Since then, the Diels-Alder reactions of tropolone, its methyl ether[4], 4-methyltropone[4], and 2-phenyltropone[5] have been reported. There still remained many points to be examined regarding the configuration of these adducts.

The present series of works was at first undertaken in order to examine the unsaturation characteristics of tropone and during the course of this work, some observations were gained regarding the configuration of the addition compound, which are described herein.

It has been pointed out[6] that the unsaturation characteristics of a tropone is greater than that of a tropolone and, as anticipated, the heating of tropone with two molar equivalents of maleic anhydride in xylene for a few hours easily effected reaction, affording the adduct I, m.p. 181.5~182.5°C, in a good yield of around 95%.

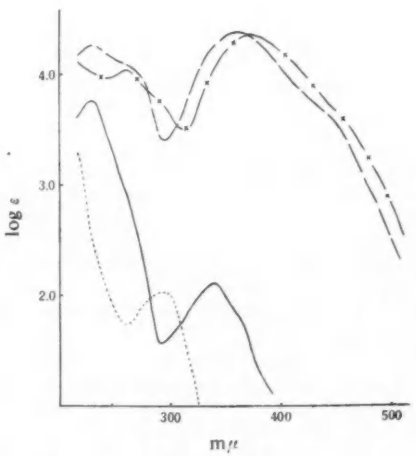


Fig. 1. —— I ----- II
-×- 2.4 DNP of V
—— 2.4 DNP of VI

Catalytic reduction of the adduct I in ethyl acetate or acetone with palladium-carbon as a catalyst, results in absorption of 2 molar equivalents of hydrogen to form a tetrahydro compound II. This shows the presence of two ethylenic linkages in I. The ultraviolet absorption spectra of I and II are shown in Fig. 1 and the spectrum of (I) exhibits a typical absorption of $\alpha\beta$-unsaturated ketone[7].

* Part IV. T. Mukai, This Bulletin, 33, 238 (1960).

** A part of this work was presented as a paper at the 6th Annual Meeting of the Chemical Society of Japan, Kyoto, April 5, 1953. Cf T. Nozoe, T. Mukai, K. Takase and T. Nagase, *Proc. Japan Acad.*, 28, 477 (1952).

1) R. D. Brown, *J. Chem. Soc.*, **1951**, 2670.

2) J. W. Cook, A. R. Gibb, R. A. Raphael and A. R. Somerville, ibid., **1951**, 503.

3) E. Sebe and C. Ôsako: Paper read before the 4th Annual Meeting of the Chemical Society of Japan in April, 1951. cf. *Proc. Japan Acad.*, **28**, 282.

4) T. Nozoe, S. Seto and T. Ikemi, *Proc. Japan Acad.*, **27**, 10 (1951).

5) T. Nozoe, T. Mukai, J. Minegishi and T. Fujisawa, *Sci. Repts. Tohoku Univ.*, Ser. I, **37**, 388 (1953).

6) T. Mukai, This Bulletin, **31**, 845 (1958).

7) R. B. Woodward, *J. Am. Chem. Soc.*, **63**, 1123 (1941); **64**, 72, 76 (1942).

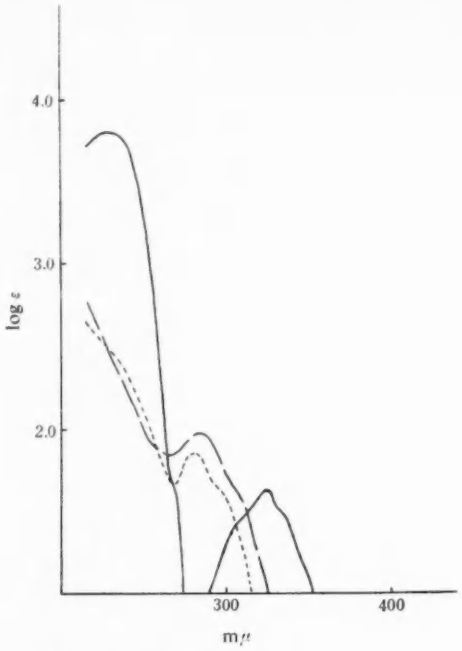


Fig. 2. —— VIII ------ X
—— XI

The infrared spectra of I and II are shown in Table I. The absorption of a carbonyl in the acid anhydride of a five-membered ring is common to I and II, but while the absorption of an $\alpha\beta$-unsaturated ketone appears at 1665 cm^{-1} in the spectrum of I, absorption of a saturated ketone is present at 1700 cm^{-1} in that of II.

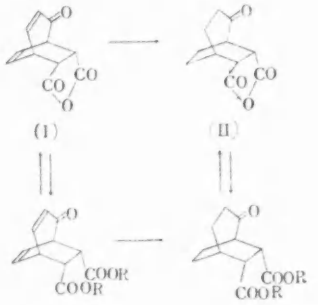
The presence of an $\alpha\beta$-unsaturated ketone in I not only offers an important clue to the structure of I but also proves the addition occurring at C-2 and C-5 positions in the tropone ring.

Treatment of the adduct I with hot water or sodium bicarbonate solution results in facile hydrolysis and a dibasic acid III is formed. The acid III is known to be dibasic from its analytical values and titration, as well as from its infrared spectrum (Table I) which shows the presence of two carbonyls.

TABLE I.

	ν_{OH} cm^{-1}	$\nu_{C=O}$ cm^{-1}	$\nu_{C=C}$ cm^{-1}
I		1840, 1780	1665, 1625
II		1840, 1775, 1700	
III	3250, 2960	1740, 1710	1650, 1625
V		1730	1670, 1633
VIII	3320	1770, 1730	1675
XI		1780, 1726, 1706	

Determination of the melting point of III indicates decomposition at 167.5°C but solidification of the melt and reexamination of the melting point shows it melting at 182.5°C, the same as that of I. This is considered to be due to the facile dehydration of III to form the anhydride.



III, R=H IV, R=H
V, R=CH_3 VI, R=CH_3

Actually, drying of III at 100°C under a reduced pressure results in the formation of I and such a fact naturally indicates that the two carboxyls are in *cis* configuration.

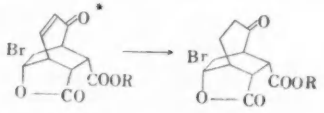
The tetrahydro compound II is also easily hydrolyzed and forms a dibasic acid (IV) which undergoes facile dehydration as in the case of III. Catalytic reduction of the unsaturated dibasic acid III in acetone, with palladium-carbon as a catalyst, results in absorption of two molar equivalents of hydrogen to form IV. Methylation of these two dibasic acids (III and IV) with diazomethane affords respective dimethyl esters (V and VI). It has been found, however, that reaction of III and excess of diazomethane also affords a compound (VII) of $C_{14}H_{16}O_5N_2$, besides V. In general, it is known that a reaction of $\alpha\beta$-unsaturated ketone and diazomethane results in formation of a pyrazole ring[8] and VII is considered to be this kind of product.

Both the methyl esters (V and VI) easily form crystalline 2, 4-dinitrophenylhydrazones whose ultraviolet absorption spectra are indicated in Fig. 1. The absorption of the 2, 4-dinitrophenylhydrazone of V shows a shift of 13 mμ to the longer wavelength region than that of VI and this fact indicates the presence of $\alpha\beta$-unsaturated ketone in V.[9] Catalytic reduction of V in ethyl acetate over palladium-carbon gives an oily product whose derivation to 2,4-dinitrophenylhydrazone proved that the oily product is VI.

8) H. Gilman, "Org. Chem.", Vol. IV, John Wiley & Sons., Inc., London, (1953), p. 780.
9) E. A. Braude and E. R. H. Jones, *J. Chem. Soc.*, 1945, 498.

Bromination of the adduct I in aqueous solution by the method of Alder-Stein[10] results in addition to the double bond and a bromolactone VIII is obtained in a good yield. Titration of VIII shows it to be a monobasic acid and the presence of a carboxyl (3320 and 1730 cm^{-1}) and a γ-lactone (1770 cm^{-1}) is also indicated from the infrared spectral values listed in Table I. In addition, there is an absorption of $\alpha\beta$-unsaturated carbonyl at 1675 cm^{-1} in the spectrum of VIII. This fact is also certain from the ultraviolet spectrum of V shown in Fig. 2. These facts indicate that the addition of BrOH occurs at the non-conjugated double bond of the two ethylenic linkages present in I. It is therefore concluded that VIII has the configuration shown in Fig. 2 and that the structure of I to VI is in *endo-cis* form.

The unsaturated bromo-lactone VIII forms a monomethyl ester IX on methylation with diazomethane and IX easily forms a 2,4-dinitrophenylhydrazone. The ultraviolet spectrum of 2,4-dinitrophenylhydrazone of IX exhibits absorption maximum at 371 mμ and



VIII, R=H X, R=H
IX, R=CH_3 XI, R=CH_3

this corresponds to that of 2,4-dinitrophenylhydrazone of an $\alpha\beta$-unsaturated ketone.

Catalytic reduction of VIII over palladium-carbon in ethyl acetate results in absorption of one mole equivalent of hydrogen and saturated bromo-lactone (X) is produced. The structure of X was confirmed by the derivation of a monomethyl ester (XI) by reaction with diazomethane since it failed to undergo crystallization.

Experimental

Diels-Alder Reaction of Tropone and Maleic Anhydride; Formation of Tropone-Maleic Anhydride Adduct (I).—A mixture of 8.36 g. (0.079 mol.) of tropone and 15.70 g. (0.16 mol.) of maleic anhydride in 60 cc. of dry xylene was refluxed in an oil bath for 2 hr. by which pale yellow crystals began to be deposited on the wall of the vessel. After refluxing for 4.5 hr, the mixture was allowed to stand overnight in a cool place and 16.6 g. (95%) of pale yellow crystals, m. p. 175~179°C, that precipitated out were collected by filtration. Several recrystallizations from acetone afforded I as colorless prismatic crystals, m. p. 181.5~182.5°C.

Found: C, 64.77; H, 4.14%. Cald. for $C_{11}H_8O_4$: C, 64.70; H, 3.95%.

I is easily soluble in acetone, ethyl acetate, ethanol, methanol and chloroform, and sparingly soluble in ether, petroleum ether, carbon tetrachloride and benzene.

Catalytic Reduction of I; Formation of the Tetrahydro Compound (II) of the Adduct.—A solution of 500 mg. of I dissolved in 20 cc. of ethyl acetate, to which 20 mg. of 5% palladium-carbon had been added, was submitted to catalytic reduction at ordinary pressure. After absorption of 2.05 mole equivalent (109 cc.) of hydrogen, the product (500 mg.) was obtained as colorless crystals of m. p. 168~181°C and several recrystallizations from acetone-ether mixture afforded II as colorless leaflets, m. p. 191~192°C.

Found: C, 63.67; H, 5.39%. Calcd. for $C_{11}H_{12}O_4$: C, 63.45; H, 5.81%.

The use of acetone in place of ethyl acetate as a solvent in this reduction also results in absorption of two mole equivalent of hydrogen and the same product is obtained.

Hydrolysis of I; Formation of the Dicarboxylic Acid (III) of the Adduct.—*a) Hydrolysis with Water:* A solution of 1 g. of I dissolved in 30 cc. of hot water was allowed to stand overnight and 530 mg. of a colorless prismatic crystals, m. p. 162.5°C (decomp.), was obtained. Concentration of its filtrate on a water bath afforded a second crop of colorless crystals, m. p. 160°C (decomp.). The two kinds of crystals were combined and washed with acetone, affording (III) as colorless crystals, m. p. 167.5°C.

Found: C, 59.34; H, 4.35%. Calcd. for $C_{11}H_{10}O_5$: C, 59.46; H, 4.54%.

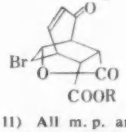
b) Hydrolysis with Sodium Bicarbonate Solution: To a solution of 420 mg. (0.005 mol.) of sodium bicarbonate dissolved in 4 cc. of water, 500 mg. of I was added and warmed on a water bath to dissolve all the crystals. The reaction mixture was concentrated acidified with hydrochloric acid while cooling with water and 400 mg. of colorless microneedles, m. p. 164~165°C (decomp.). The crystals were collected and washed several times with hot acetone, giving III as colorless microneedles, m. p. 167.5°C.

Ultraviolet spectrum: λ_{max}^{MeOH} mμ (log ε): 228 (3.87), 337 (2.49).

The acid III obtained by both the foregoing methods (a) and (b) showed no depression of the melting point on admixture with I. The melted III, upon solidification, melted at 181.5~182.5°C, same as I.

Dehydration of III; Formation of I.—The Acid III obtained from the both methods (a) and (b) as above was each kept in a vacuum (10 mmHg) desiccator over phosphorus pentoxide

10) K. Alder und G. Stein, *Anal. Chem.*, **514**, 4 (1934).



In this case there are two possible structures to VIII. Detailed reports on the structure of VIII will be made elsewhere.

11) All m. p. are uncorrected. The microanalyses were carried out by Miss Ayako Iwanaga of this laboratory to whom the authors are indebted.

at 100°C for 2 days. Determination of the melting point of resultant product showed formation of 45 mg. of a product of m. p. 181.5~182.5°C from 50 mg. of III. No depression of the melting point was observed on admixture with I.

Found: C, 64.59; H, 4.00%. Calcd. for $C_{11}H_8O_4$: C, 64.70; H, 3.95%.

Hydrolysis of II; Formation of Dicarboxylic Acid (IV) of the Tetrahydro Derivative of the Adduct.—A solution of 150 mg. of II dissolved in 3 cc. of hot water was allowed to cool and colorless crystals that separated out were collected to 100 mg. of IV as colorless microneedles, m. p. 173°C (decomp.).

Found: C, 58.33; H, 5.80%. Calcd. for $C_{11}H_{14}O_5$: C, 58.40; H, 6.24%.

The filtrate was concentrated at room temperature under a reduced pressure and a second crop of crystals was obtained as 80 mg. of colorless crystals, m. p. 169~170°C (decomp.).

Ultraviolet spectrum: λ_{max}^{MeOH} mμ (log ε): 285 (1.84).

Admixture of IV and II showed no depression of the melting point. Determination of the melting point after solidification of melted II showed m. p. 191~192°C.

Dehydration of IV; Formation of the Tetrahydro Compound (II) of the Adduct.—Thirty milligrams of IV was kept over phosphorus pentoxide in a vacuum (10 mmHg) desiccator at 100°C for two days and 28 mg. of the resultant product melted at 171~172°C, alone and in admixture with II.

Found: C, 58.43; H, 6.30%. Calcd. for $C_{11}H_{12}O_4$: C, 58.40; H, 6.24%.

Catalytic Reduction of III; Formation of IV.—A solution of 500 mg. of III dissolved in 70 cc. of acetone, added with 20 mg. of 5% palladium-carbon, was submitted to catalytic reduction at ordinary pressure and 2.08 mole equivalents of hydrogen (105 cc.) was absorbed. The catalyst was removed by filtration and the filtrate was concentrated, affording 460 mg. of colorless crystals, m. p. 154~155.5°C. The crystals were washed several times with warm acetone and (IV) was obtained as colorless crystals of m. p. 173°C (decomp.).

Found: C, 58.30; H, 5.93%. Calcd. for $C_{11}H_{14}O_5$: C, 58.40; H, 6.24%.

The compound (IV) obtained by this means easily lost one mole of water to form II, as was the case with IV obtained by the hydrolysis of II.

Raction of (III) and Diazomethane.—*a) Formation of the Dimethyl ester (V) of III*: To a suspension of 220 mg. (ca. 0.001 mol.) of finely pulverized III in 10 cc. of ethyl acetate, 3.6 cc. (0.002 mol.) of ether solution of diazomethane was added dropwise at room temperature by which a vigorous reaction set in and III dissolved almost completely. The solution was evaporated on a water bath, the concentrated solution was cooled in ice, and 20 mg. of colorless crystals VII, m. p. 212°C, separated out.

Concentration of the mother liquor left 200 mg. of yellow oil which solidified gradually upon standing. Recrystallization from a small volume of methanol afforded V as yellow microneedles, m. p. 83~88°C.

Found: C, 62.42; H, 5.52%. Calcd. for $C_{13}H_{14}O_5$: C, 62.39; H, 5.64%.

Ultraviolet spectrum: λ_{max}^{MeOH} mμ (log ε): 227 (3.90), 337(2.40).

b) Formation of VII: To a suspension of 300 mg. (0.0014 mol.) of III in 20 cc. of ethyl acetate, 9 cc. (0.0042 mol.) of ether solution of diazomethane was added in drops at room temperature by which III dissolved to form a uniform yellow solution. On standing, colorless crystals separated out which were collected and recrystallized from water to 340 mg. of colorless needles (VII), m. p. 212°C (decomp.).

Found: C, 57.19; H, 5.13%. Calcd. for $C_{13}H_{14}O_5 \cdot CH_2N_2$: C, 57.53; H, 5.52%.

Ultraviolet spectrum: λ_{max}^{MeOH} mμ (log ε): 235 (3.36), 307(3.68).

c) 2,4-Dinitrophenylhydrazone of V: A solution (conc. sulfuric acid, 1.3 cc., methanol, 6.5 cc.) of 250 mg. (0.0013 mol.) of 2,4-dinitrophenylhydrazine was added to a solution of 250 mg. (0.001 mol.) of V dissolved in 2.5 cc. of methanol by which a fine orange powder precipitated out immediately. The mixture was stirred and allowed to stand overnight at room temperature. The solid was collected and recrystallized several times from acetone-methanol mixture to 230 mg. of VII as lustrous orange powder, m. p. 221.5~222°C.

Found: C, 52.80; H, 3.96; N, 13.07%. Calcd. for $C_{19}H_{15}O_8N_4$: C, 53.02; H, 4.22; N, 13.02.

Formation of the Dimethyl Ester (VI) of IV.—(a) The calculated amount (0.004 mol.) of the ether solution of diazomethane was added dropwise at room temperature into a solution of 450 mg. (0.002 mol.) of IV dissolved in 20 cc. of ethyl acetate by which a vigorous effervescence set in. After standing at room temperature for a few hours, the solvent was distilled off and left 480 mg. of pale yellow oil. This oily residue was dissolved in acetone and purified by passage through alumina layer, from which 420 mg. of pale yellow oil (VI) was obtained.

Ultraviolet spectrum: λ_{max}^{MeOH} mμ (log ε): 284(1.19).

b) Catalytic reduction of 50 mg. of V dissolved in 10 cc. of ethyl acetate, with 5% palladium-carbon as a catalyst, resulted in absorption of 2 mol. equivalents (5 cc., calcd. 4.5 cc.) and 50 mg. of pale yellow oil (VI) was obtained.

Ultraviolet spectrum: λ_{max}^{MeOH} mμ (log ε): 284(1.20).

c) Formation of 2,4-Dinitrophenylhydrazone of VI: The hydrazone was prepared as in the case of V and recrystallized from methanol to fine orange needles, m. p. 89°C (decomp.).

Found: C, 52.50; H, 4.90; N, 12.87%. Calcd. for $C_{19}H_{22}O_8N_4$: C, 52.53; H, 5.10; N, 12.90%.

Formation of Bromolactone-monocarboxylic Acid (VIII).—To a solution of 500 mg. (0.0025 mol.) of I dissolved in 30 cc. of water, 500 mg. (0.0031 mol.) of bromine was dropped in under ice-cooling and stirring. After some time, pale yellow crystals began to precipitate out. The mixture was stirred under ice-cooling for a further 4 hr., the excess of bromine was removed by addition of sodium bisulfite, and the mixture was filtered, affording 570 mg. of fine colorless crystals, m. p. 185°C

(decomp.). After drying over phosphorus pentoxide in a vacuum desiccator, the crystals were recrystallized from acetone to 500 mg. of colorless plates (VIII), m. p. 219°C (decomp.). VIII dissolved in cold, saturated solution of sodium bicarbonate with effervescence and colors green to the Beilstein test.

Found: C, 43.81; H, 2.99%. Calcd. for $C_{11}H_9O_5Br$: C, 43.87; H, 3.01%.

Formation of the Methyl Ester (IX) of the Bromo-lactone VIII.—To a solution of 300 mg. (0.001 mol.) of VIII dissolved in 10 cc. of ethyl acetate, 1.9 cc. (0.001 mol.) of the ether solution of diazomethane was added in drops and the mixture was allowed to stand for a few days from which 150 mg. of colorless scaly crystals of m. p. 213~215°C separated out. These crystals were collected by filtration, the filtrate was concentrated, and a second crop of crystals were obtained as 150 mg. of pale yellow crystals, m. p. 178~195°C. The two crops of crystals were combined and recrystallized from acetone-methanol mixture to IX as colorless scales, m. p. 216~217°C.

Found: C, 45.69; H, 3.25%. Calcd. for $C_{12}H_{11}O_5Br$: C, 45.73; H, 3.52%.

Ultraviolet spectrum: λ_{max}^{MeOH} mμ (log ε): 232 (3.82), 320(2.37).

2,4-Dinitrophenylhydrazone of IX: A mixture of 200 mg. of IX dissolved in 20 cc. of methanol and methanolic solution (3.5 cc. of methanol and 0.9 cc. of conc. sulfuric acid) of 150 mg. of 2,4-dinitrophenylhydrazine was refluxed on a water bath for 3.5 hr. Concentration of this solution until only a part of the solvent remained left 220 mg. of dark brown crystals which were collected and recrystallized repeatedly from methanol to dark brown, fine needles, m. p. 146~147°C (decomp.).

Found: C, 43.50; H, 3.00; N, 11.02%. Calcd. for $C_{18}H_{15}O_8N_4Br$; C, 43.65; H, 3.05; N, 11.31%.

Ultraviolet spectrum: λ_{max}^{MeOH} mμ (log ε): 255 (4.13), 371(4.43).

Catalytic Reduction of VIII; Formation of Dihydro-bromo-lactone (X) and Methyl Ester (XI). of X,—A solution of 240 mg. of VIII dissolved in 20 cc. of ethyl acetate, to which 20 mg. of palladium-carbon was added, was submitted to reduction at ordinary pressure and 0.95 mole equivalent (24.5 cc.) of hydrogen was absorbed. The product obtained therefrom was 220 mg. of yellowish brown oil (X) which failed to crystallize.

A calculated amount of ether solution of diazomethane was added to a solution of 120 mg. of above oil dissolved in 1 cc. of ethyl acetate and the mixture was allowed to stand overnight, from which 100 mg. of colorless crystals precipitated out. This was recrystallized from methanol-acetone mixture to colorless fine powder (XI), m. p. 167.5°C.

Found: C, 45.46; H, 4.02%. Calcd. for $C_{12}H_{13}O_5Br$: C, 45.44; H, 4.13%.

Chemical Institute
Faculty of Science
Tohoku University
Katahira-cho, Sendai

---

*Theory of the Influence of the Dielectric Constant on the Rate of Reaction in Solution with Application to Enzyme Reactions. I. Development of the Theory and its Application to Some Simple Systems**

By Keitaro HIROMI

(Received December 18, 1959)

The rate of the reaction in solution involving ions or dipolar molecules is generally influenced by the dielectric constant of the medium, and this effect, termed the solvent effect, has been quantitatively interpreted by various authors in conformity to the electrostatic theories[1~5]. There is, however, marked inconsistency between the prediction of Amis-Jaffé's theory and that of Laidler-Eyring's for reactions between a negative ion and a dipolar molecule[2]. Recently Laidler and Landskroener developed a more general theory of the solvent effect[6] based on Kirkwood's model for a spherical solute molecule having an arbitrary charge

* Presented in part at the 11th Annual Meeting of Chem. Soc. Japan, Tokyo, April 1958, and at the Symposium on Enzyme Chem., Sapporo, July, 1958.

1) S. Glasstone, K. J. Laidler and H. Eyring, "The Theory of Rate Processes", McGraw-Hill, New York (1954), p. 400.

2) E. S. Amis, "Kinetics of Chemical Change in Solution", MacMillan Co., New York (1949), p. 153.

3) E. S. Amis, *J. Chem. Educ.*, **28**, 635 (1951); **29**, 337 (1952); **30**, 351 (1953).

4) E. A. Moelwyn-Hughes, "The Kinetics of Reactions in Solution", Oxford University Press, London (1947), p. 90, 120, 202.

5) G. Scatchard, *J. Am. Chem. Soc.*, **52**, 52 (1930).

6) K. J. Laidler and P. A. Landskroener, *Trans. Faraday Soc.*, **52**, 200 (1956).

distribution[7,8], considering the microscopic distribution of charges in the reactants and the activated complexes, which had not been taken into account explicitly in the previous treatments[1-5]. Their theory was applied to the interpretation of the effect of solvent on the rates of the acid and base hydrolyses of esters, amides and anilides, and the results were considered to be successful[6].

The influence of the dielectric constant upon the rates of several enzyme reactions have been treated quantitatively on the basis of simple models[9,10], which seem evidently inadequate to a large complex ion such as protein. The Kirkwood's model may be a suitable one for the treatment of the electrostatic properties of protein molecules[11-13]. Although Laidler-Landskroener's theory is based on this model, it is not applicable to enzyme reactions, because of the approximation made in its derivation, which is not adequate for the large molecules with their charges near the surface of the sphere.

In this paper, a theory of the solvent effect which is applicable to enzyme reactions as well as to reactions involving small ions or molecules is developed, based on the Kirkwood's model. The theory is applied to the interpretation of the effect of solvent on the dissociation constants of carboxylic acids and on the rate constants of unimolecular and bimolecular reactions of simple organic compounds. Application of the theory to enzyme reactions will be treated in the following paper.

Theoretical

Kirkwood[7] obtained an expression for the work of charging, or the electrostatic free energy, of a spherical solute molecule of radius b, having a low dielectric constant D_i, within which are situated M discrete point charges $e_1 \cdots e_M$, and which is immersed in a solvent of dielectric constant D, as illustrated in Fig. 1. For zero ionic strength, the work of charging, W_0, with respect to the energy zero for the completely discharged state, is expressed as follows[7,8,12,14]:

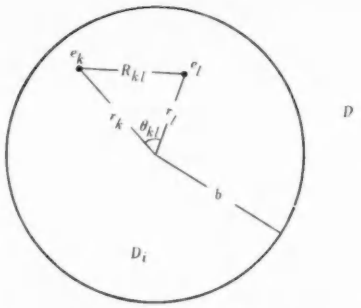


Fig. 1. Kirkwood's model for a spherical solute species having an arbitrary charge distribution.

$$W_0 = \frac{1}{2D_i}\sum_{k=1}^{M}\sum_{l=1}^{M} e_k e_l / R_{kl} + \frac{1}{2}\left(\frac{1}{D}-\frac{1}{D_i}\right)\sum_{n=0}^{\infty}\frac{(n+1)Q_n}{b^{2n+1}[(n+1)+nD_i/D]} \quad (1)$$

$$Q_n = \sum_{k=1}^{M}\sum_{l=1}^{M} e_k e_l r_k^n r_l^n P_n(\cos\theta_{kl}) \quad (2)$$

where R_{kl} is the distance between the charges e_k and e_l, r_k and r_l the distances of charges e_k and e_l from the center of the sphere, θ_{kl} the angle between r_k and r_l, and $P_n(\cos\theta_{kl})$ is the Legendre polynomial. Q_0 is simply the square of the total net charge of the ion, and Q_1 is equal to the square of the dipole moment of the molecule. The succeeding Q_n's involve the higher multipole moments, and depend not only upon the relative position of the charges, but also upon the distance of the charges from the surface of the sphere.

When the charges are situated not so near to the surface of the sphere, i. e., r_k/b is not so close to unity, the summation in Eq. 1 converges rapidly, and the terms beyond $n=1$ can be dropped without serious error. This is the case in Laidler and Landskroener's treatment[6]. However, this approximation is not a reasonable one, especially when b is large and the charges are situated near the surface, as is the case of a large complex ion such as protein. Now it is necessary to take the sum of the infinite series in Eq. 1. Since $D_i \ll D$ (D_i was taken by Kirkwood to be 2), in a similar manner to Kirkwood's treatment[8], the sum is obtained by expanding the second term of Eq. 1 in increasing powers of D_i/D, and dropping all terms

7) J. G. Kirkwood, *J. Chem. Phys.*, **2**, 351 (1934).
8) J. G. Kirkwood and F. H. Westheimer, ibid., **6**, 506 (1938).
9) Y. Ogura et al., *Symposia on Enzyme Chem.* (*Kôso Kagaku Shimpojiumu*), **5**, 28 (1950); Y. Tonomura, *J. Japan. Biochem. Soc.*, (*Seikagaku*), **25**, 157 (1953).
10) M. L. Barnard and K. J. Laidler, *J. Am. Chem. Soc.*, **74**, 6099 (1952); K. J. Laidler and M. C. Ethier, *Arch. Biochem. Biophys.*, **44**, 338 (1953).
11) T. L. Hill, *J. Phys. Chem.*, **60**, 253 (1956); *J. Am. Chem. Soc.*, **78**, 3330 (1956); ibid., **78**, 5527 (1956).
12) C. Tanford and J. G. Kirkwood, ibid. **79**, 5333 (1957).
13) C. Tanford, ibid., **79**, 5340 (1957).
14) The first term in Eq. 1 represents the work of charging in an unbounded medium of dielectric constant D_i, and is independent of the solvent. It contains self-energy terms (i. e., $k=l$), which are infinite because of the assumption of point charges. But this term does not appear in the final expression for the activity coefficient (Eq. 8 below). The second term represents the mutual electrostatic energy of the molecule and the solvent medium, and is the important term in the present treatment.

higher than the first. Finally we have[15]

$$W_0=\frac{1}{2D_i}\left[\sum_{k=1}^{M}\sum_{l=1}^{M}\frac{e_ke_l}{R_{kl}}-\frac{1}{b}\left(\sum_{k=1}^{M}e_k^2f'_{kk}+\sum_{k=1}^{M}\sum_{l\neq k}e_ke_lg'_{kl}\right)\right]+\frac{1}{2Db}\left[\sum_{k=1}^{M}e_k^2f_{kk}+\sum_{k=1}^{M}\sum_{l\neq k}e_ke_lg_{kl}\right] \quad (3)$$

where

$$f_{kk}=\frac{2}{1-x_{kk}}+\frac{1}{x_{kk}}\ln(1-x_{kk})$$
$$g_{kl}=\frac{2}{(1-2\alpha_{kl}x_{kl}+x_{kl}^2)^{1/2}}+\frac{1}{x_{kl}}\ln\frac{1-\alpha_{kl}}{(1-2\alpha_{kl}x_{kl}+x_{kl}^2)^{1/2}+x_{kl}-\alpha_{kl}} \quad (4)$$

$$f'_{kk}=\frac{1}{1-x_{kk}}$$
$$g'_{kl}=\frac{1}{(1-2\alpha_{kl}x_{kl}+x_{kl}^2)^{1/2}} \quad (5)$$

where

$$x_{kl}=r_kr_l/b^2;\quad x_{kk}=(r_k/b)^2$$
$$\alpha_{kl}=\cos\theta_{kl} \quad (6)$$

For zero ionic strength, the activity coefficient γ of a solute species in the solvent of dielectric constant D with respect to a standard solvent of dielectric constant D_0 ($D_0\gg D_i$) is related to the work of charging W_0 and W_0^0 in each solvent, respectively, in the following manner

$$\ln\gamma=(W_0-W_0^0)/kT \quad (7)$$

where k is the Boltzmann constant and T the absolute temperature. Insertion of Eq. 3 into Eq. 7 gives rise to

$$\ln\gamma=\frac{\varepsilon^2L}{2bkT}\left(\frac{1}{D}-\frac{1}{D_0}\right) \quad (8)$$

where ε is the protonic charge and L is a new dimensionless parameter defined as

$$L=\sum_{k=1}^{M}\xi_k^2f_{kk}+\sum_{k=1}^{M}\sum_{l\neq k}\xi_k\xi_lg_{kl} \quad (9)$$

where ξ_k is the valence of the charge e_k, i. e., $\xi_k=e_k/\varepsilon$. The value of L is determined by the charge configuration of the species, i. e., L is a function of ξ's, r's, θ's of the charges and the radius b. Thus L will be conveniently termed "charge configuration function". L/b represents the susceptibility of the electrostatic free energy of the solute species to the change in dielectric constant of the medium.

By the use of this new parameter L, a relationship between the rate or the equilibrium constant of a chemical change in solution and the dielectric constant of the solvent can readily be obtained. For example, let us consider the reaction

$$A+B\rightleftarrows X_*\rightarrow \text{Products}$$

where the reactants A, B and the activated complex X_* may be either ions or molecules with particular charge configuration. Let k_0 denote the rate constant in the standard solvent of dielectric constant D_0 ($D_0\gg D_i$), then the rate constant k in a solvent of dielectric constant D is[1]

$$k=k_0\gamma_A\gamma_B/\gamma_* \quad (10)$$

where the γ's are the activity coefficients of the corresponding species. Using Eq. 8 we have

$$\ln k=\ln k_0+\frac{\varepsilon^2}{2kT}\left(\frac{1}{D}-\frac{1}{D_0}\right)\left(\frac{L_A}{b_A}+\frac{L_B}{b_B}-\frac{L_*}{b_*}\right) \quad (11)$$

where L's and b's are the charge configuration functions and the radii of the species, respectively. This equation indicates that the plot of $\log k$ against $1/D$ gives a straight line and that the slope is

$$\frac{d\log k}{d(1/D)}=\frac{\varepsilon^2}{2.303\,(2kT)}\left(\frac{L_A}{b_A}+\frac{L_B}{b_B}-\frac{L_*}{b_*}\right) \quad (12)$$

Thus the slope is determined by the charge configuration functions and the radii of the species.

In a similar manner, the relationship between the equilibrium constant K and the dielectric constant of the medium, D, can be obtained. For the equilibrium

Table I. Values of f for various values of x

x	f	x	f
0	1.000	0.925	23.866
0.05	1.079	0.930	25.712
0.1	1.169	0.935	27.846
0.2	1.384	0.940	30.340
0.3	1.668	0.945	33.294
0.4	2.056	0.950	36.847
0.5	2.614	0.955	41.197
0.6	3.473	0.960	46.647
0.7	4.947	0.965	53.669
0.8	7.988	0.970	63.052
0.83	9.630	0.975	76.217
0.85	11.101	0.980	96.008
0.88	14.257	0.985	129.070
0.90	17.442	0.990	195.348
0.91	19.576	0.995	394.675
0.92	22.255	1	∞

15) Eq. 2 can be written as

$$Q_n=\sum_{k=1}^{M}e_k^2r_k^{2n}+\sum_{k=1}^{M}\sum_{l\neq k}e_ke_lr_k^nr_l^nP_n(\cos\theta_{kl})$$

When this is inserted into Eq. 1 and the summation is performed, the first term gives the f and f' terms and the second term gives the g and g' terms in Eq. 3.

TABLE II. VALUES OF g FOR

$\alpha \backslash x$	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.85	0.90
−1	0.865	0.755	0.664	0.587	0.522	0.467	0.418	0.376	0.357	0.339
−0.9	0.876	0.772	0.683	0.607	0.542	0.485	0.436	0.392	0.372	0.354
−0.8	0.887	0.789	0.703	0.628	0.563	0.505	0.454	0.409	0.389	0.370
−0.7	0.899	0.807	0.725	0.651	0.585	0.527	0.475	0.428	0.407	0.387
−0.6	0.911	0.827	0.748	0.676	0.610	0.551	0.497	0.449	0.427	0.406
−0.5	0.924	0.847	0.773	0.702	0.637	0.577	0.522	0.473	0.450	0.428
−0.4	0.936	0.868	0.799	0.732	0.667	0.607	0.551	0.499	0.475	0.452
−0.3	0.950	0.891	0.828	0.764	0.700	0.639	0.582	0.528	0.503	0.479
−0.2	0.963	0.915	0.859	0.799	0.737	0.677	0.618	0.562	0.536	0.510
−0.1	0.977	0.941	0.893	0.838	0.779	0.719	0.659	0.601	0.573	0.546
0	0.992	0.968	0.930	0.882	0.826	0.767	0.706	0.646	0.617	0.588
0.1	1.007	0.997	0.971	0.931	0.881	0.823	0.762	0.699	0.668	0.637
0.2	1.022	1.028	1.015	0.986	0.943	0.889	0.828	0.763	0.730	0.697
0.3	1.038	1.061	1.065	1.050	1.017	0.968	0.909	0.842	0.807	0.771
0.4	1.055	1.096	1.120	1.123	1.104	1.065	1.009	0.941	0.904	0.865
0.5	1.072	1.135	1.182	1.209	1.211	1.187	1.138	1.071	1.032	0.990
0.6	1.090	1.176	1.253	1.311	1.344	1.345	1.312	1.250	1.209	1.163
0.7	1.108	1.222	1.333	1.435	1.515	1.561	1.562	1.515	1.475	1.426
0.8	1.128	1.271	1.427	1.590	1.747	1.877	1.956	1.960	1.931	1.883
0.9	1.148	1.325	1.537	1.789	2.080	2.398	2.702	2.916	2.961	2.953
0.92	1.152	1.336	1.561	1.836	2.166	2.547	2.945	3.270	3.363	3.388
0.94	1.156	1.348	1.586	1.886	2.260	2.720	3.248	3.750	3.930	4.021
0.96	1.160	1.360	1.613	1.939	2.365	2.925	3.640	4.450	4.810	5.058
0.98	1.164	1.372	1.640	1.995	2.482	3.171	4.172	5.597	6.433	7.208
0.99	1.166	1.378	1.654	2.025	2.546	3.313	4.518	6.527	7.980	9.679
0.995	1.168	1.381	1.661	2.041	2.579	3.391	4.720	7.163	9.210	12.123
0.997	1.168	1.382	1.664	2.047	2.593	3.423	4.808	7.466	9.859	13.671
0.998	1.168	1.383	1.665	2.050	2.600	3.439	4.853	7.630	10.231	14.671
0.999	1.168	1.384	1.667	2.053	2.607	3.456	4.899	7.804	10.643	15.895
1	1.169	1.384	1.668	2.056	2.614	3.473	4.947	7.988	11.101	17.442

* When x is zero, the value of g is equal to unity for any value of α.

$$A+B \rightleftharpoons C+D$$

the final expression is

$$\frac{d \log K}{d(1/D)} = \frac{\varepsilon^2}{2.303(2kT)}\left[\left(\frac{L_A}{b_A}+\frac{L_B}{b_B}\right)-\left(\frac{L_C}{b_C}+\frac{L_D}{b_D}\right)\right] \quad (13)$$

Assuming that the molecule or ion is a sphere within which each point charge is located at the center of each constituting atom, the value of L of the molecule or ion with known charge configuration can be calculated using Eqs. 4, 6 and 9. For convenience of calculation, the values of f and g for various values of x and α are tabulated in Tables I and II. For $x=0$, f and g are both equal to unity. The value of g for $\alpha=1$ is equal to that of f for the corresponding value of x. The value of L is dependent on the position of the center to be chosen. It may be most reasonable to choose the center at the geometrical center of the

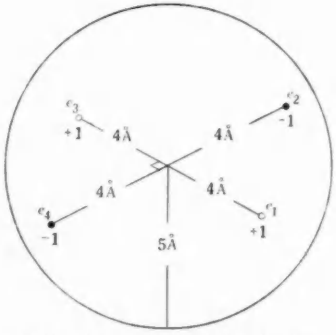


Fig. 2. A hypothetical dipolar ion. Two sets of charges, unit positive charges, e_1 anp e_3, and unit negative charges e_2 and e_4, are situated at equal distance 4 Å from the center, on the two diameters which are perpendicular to each other. The radius of the sphere is 5 Å.

VARIOUS VALUES OF x AND α*

0.91	0.92	0.93	0.94	0.95	0.96	0.97	0.98	0.99	1	x/α
0.336	0.333	0.329	0.326	0.323	0.319	0.316	0.313	0.310	0.307	−1
0.350	0.347	0.343	0.340	0.336	0.333	0.330	0.326	0.323	0.320	−0.9
0.366	0.362	0.359	0.355	0.351	0.348	0.344	0.341	0.338	0.334	−0.8
0.383	0.379	0.376	0.372	0.368	0.364	0.361	0.357	0.354	0.350	−0.7
0.402	0.398	0.394	0.390	0.386	0.383	0.379	0.375	0.371	0.368	−0.6
0.424	0.419	0.415	0.411	0.407	0.403	0.399	0.395	0.391	0.387	−0.5
0.447	0.443	0.439	0.434	0.430	0.426	0.422	0.417	0.413	0.409	−0.4
0.474	0.470	0.465	0.460	0.456	0.451	0.447	0.442	0.438	0.434	−0.3
0.505	0.500	0.495	0.490	0.486	0.481	0.477	0.471	0.467	0.462	−0.2
0.541	0.535	0.530	0.525	0.520	0.515	0.510	0.505	0.500	0.495	−0.1
0.582	0.576	0.571	0.565	0.560	0.554	0.549	0.544	0.538	0.533	0
0.631	0.625	0.619	0.613	0.607	0.602	0.596	0.590	0.584	0.578	0.1
0.691	0.684	0.678	0.671	0.665	0.658	0.652	0.646	0.639	0.633	0.2
0.764	0.757	0.750	0.743	0.736	0.729	0.722	0.714	0.708	0.701	0.3
0.858	0.850	0.842	0.834	0.826	0.818	0.810	0.803	0.795	0.787	0.4
0.981	0.972	0.964	0.955	0.946	0.937	0.928	0.919	0.910	0.901	0.5
1.154	1.144	1.134	1.124	1.114	1.104	1.093	1.083	1.072	1.062	0.6
1.415	1.404	1.393	1.381	1.369	1.357	1.344	1.332	1.319	1.306	0.7
1.871	1.858	1.845	1.831	1.817	1.802	1.786	1.770	1.753	1.736	0.8
2.944	2.933	2.921	2.905	2.888	2.869	2.848	2.824	2.799	2.772	0.9
3.383	3.376	3.365	3.352	3.335	3.315	3.293	3.267	3.239	3.208	0.92
4.026	4.027	4.022	4.013	3.999	3.981	3.957	3.929	3.897	3.860	0.94
5.088	5.111	5.125	5.131	5.128	5.116	5.095	5.066	5.029	4.983	0.96
7.335	7 448	7.544	7.622	7.678	7.711	7.720	7.706	7.666	7.602	0.98
10.018	10.343	10.645	10.917	11.149	11.330	11.454	11.514	11.504	11.425	0.99
12.808	13.512	14.222	14.915	15.565	16.137	16.594	16.900	17.026	16.955	0.995
14.665	15.740	16.884	18.074	19.270	20.407	21.399	22.143	22.542	22.531	0.997
15.909	17.295	18.832	20.513	22.300	24.113	25.811	27.192	28.027	28.138	0.998
17.486	19.352	21.549	24.134	27.148	30.572	34.241	37.713	40.216	40.899	0.999
19.576	22.255	25.712	30.340	36.847	46.647	63.052	96.008	195.348	∞	1

molecule or ion concerned. As an example, L of a hypothetical dipolar ion such as that depicted in Fig. 2 is calculated to be 11.8. The procedure of the calculation is illustrated in Appendix I.

For a spherical ion including a charge Ze at its center, L is simply Z^2 (since $x=0$ and f is unity). Thus for a reaction between such unipolar ions Eq. 12 reduces to the familiar expression of Laidler-Eyring for reactions between ions[1].

For the purpose of comparison, Laidler-Landskroener's equation[6], which corresponds to Eq. 12, is cited below;

$$\frac{\mathrm{d}\log k}{\mathrm{d}(1/D)}=\frac{\varepsilon^2}{2.303(2kT)}\left[\left(\frac{Z_A{}^2}{b_A}+\frac{Z_B{}^2}{b_B}-\frac{(Z_A+Z_B)^2}{b_*}\right)+\frac{3}{2}\left(\frac{G_A}{b_A{}^3}+\frac{G_B}{b_B{}^3}-\frac{G_*}{b_*{}^3}\right)\right] \quad (14)$$

where b and Z represent the radius and the net charge of the species, respectively, and G is a quantity determined by the given charge distribution in the species, which is proportional to the square of the dipole moment.

Application of the Theory to Some Simple Systems

Before applying the theory to enzyme reactions, it is desirable to test its applicability to some simple systems, for which the structure and the charge distribution in the species concerned can be visualized with considerable certainty. Three examples will be treated here; first, the dissociation equilibria of carboxylic acids; second, the S_N1 solvolysis; and third, the acid and base catalyzed hydrolyses of esters, amides and anilides.

Dissociation of Carboxylic Acids. — The dissociation constants of some carboxylic acids have been measured by several authors in

TABLE III. THE INFLUENCE OF THE DIELECTRIC CONSTANT UPON THE DISSOCIATION CONSTANTS OF CARBOXYLIC ACIDS

Acid	Temp. °C	Solvent[a]	Dissociation constant	$d(pK)/d(1/D)$	b(Å)[b] I	b(Å)[b] II	d(Å)[b] I	d(Å)[b] II
Acetic	25	W, W-E / W-D	K	118	2.97		1.33	
Benzoic	20	W, W-E / W-D	K	170	3.75		0.65	
Succinic	20	W, W-E / W-M, W-D	K_1	133	4.30	4.40	1.09	1.19
				150	4.07	4.10	0.86	0.89
			K_2	178	3.73	3.80	1.01	1.08
				200	3.48	3.50	0.76	0.78
Adipic	20	W, W-E / W-D	K_1	147		4.84 (all *trans*)		0.99
						3.95 (di-*cis*)		0.94
			K_2	164		4.88 (all *trans*)		1.03
						4.06 (di-*cis*)		1.05
Azelaic	20	W, W-E / W-M, W-D	K_1	135		6.84 (all *trans*)		1.23
						5.84 (tri-*cis*)		1.21
			K_2	140		6.98 (all *trans*)		1.37
						6.03 (tri-*cis*)		1.40

a) W: water, E: ethanol, M: methanol, D: dioxane.
b) b (or d) indicates the radius (or the depth of the charge on the oxygen atom in COO^-) of A^-, AH^- or $A^=$ corresponding to the dissociation constant K, K_1 or K_2, respectively. I: Calculated by Method I. II: Calculated by Method II.

various mixed aqueous solvents. The apparent dissociation constant K_a of an acid AH ($K_a = [H^+][A^-]/[AH]$) in various mixed solvents having equal dielectric constants does not give the same value in different media having different water concentrations[16,17]. Yasuda[18] proposed that the dissociation constant K of an acid in aqueous mixed solvents should be given by the following equation:

$$K = [A^-][H_3O^+]/[AH][H_2O]$$

corresponding to the equilibrium

$$AH + H_2O \rightleftarrows H_3O^+ + A^-$$

He found that the plot of pK, instead of pK_a, of an acid against $1/D$ becomes a single straight line, irrespective of the sort of mixed solvent used, unless $1/D$ exceeds ca. 0.02[18]. Therefore, it may be reasonable to conclude that the variation in pK is essentially due to the change in the dielectric constant of the medium, at least in the range stated above. The slope $d(pK)/d(1/D)$ is written as follows:

$$-\frac{d(pK)}{d(1/D)} = \frac{\varepsilon^2}{2.303(2kT)}\left(\frac{L_{AH}}{b_{AH}} + \frac{L_{H_2O}}{b_{H_2O}} - \frac{L_{A^-}}{b_{A^-}} - \frac{L_{H_3O^+}}{b_{H_3O^+}}\right) \qquad (15)$$

Values of the slopes obtained from Yasuda's results for several carboxylic acids are summarized in Table III, where K_1 and K_2 are the first and the second dissociation constants of dibasic acids, respectively.

$$K_1 = [AH^-][H_3O^+]/[AH_2][H_2O]$$

$$K_2 = [A^=][H_3O^+]/[AH^-][H_2O]$$

Procedure of testing the theory is as follows; first, for the three species, H_2O, H_3O^+ and AH (or AH_2 for the dibasic acid), L/b's are calculated from the charge configurations and the radii of these species; insertion of these values together with the value of the slope into Eq. 15 leads to an equation for L_{A^-}/b_{A^-}; if the magnitudes and the relative positions of the charges in A^- are known, L_{A^-} and L_{A^-}/b_{A^-} are the functions of b_{A^-} only; then the equation can be solved for b_{A^-}[19]. If the theory is correct, b_{A^-} thus obtained should be of reasonable magnitude as is expected from the structure of A^-. The geometry and the charge distribution of the species which are necessary for the calculation of L's are summarized in Appendix II.

Succinic acid was treated by the following two methods. In Method I, the charge distribution in the COOH group was considered explicitly in the evaluation of L_{AH_2} and L_{AH^-} as in the case of acetic acid and benzoic acid. Since L_{AH_2} was found to be appreciably small,

16) J. C. Speakman, *J. Chem. Soc.*, **1943**, 270.
17) M. L. Dondon, *J. chim. phys.*, **54**, 290, 304 (1957).
18) M. Yasuda, This Bulletin, **32**, 429 (1959).

19) Practically b_{A^-} is obtained in the following manner. For several fixed values of b_{A^-} suitably chosen, L_{A^-}'s are calculated and L_{A^-}/b_{A^-} is plotted against b_{A^-}. By graphical interpolation b_{A^-} to be solved is readily found.

TABLE IV. THE INFLUENCE OF SOLVENT UPON THE RATE CONSTANTS OF S_N1 SOLVOLYSIS

Compound	Temp. °C	Solvent[a]	$\frac{d \log k_1}{d(1/D)}$	Model	Type	Present theory		L-L theory
						b_*(Å)	d(Å)[b]	b_*(Å)[c]
tert-Butyl chloride	25	(W-E, W-M W-A, HCOOH)	−179	I	(1)	3.12	0.48	
					(2)	3.10	0.46	1.92
				II	(1)	2.41	0.47	
					(2)	2.36	0.42	
tert-Butyl bromide	25	(W-E, W-A)	−127	I	(1)	3.50	0.64	
					(2)	3.46	0.60	2.27
				II	(1)	2.72	0.62	
					(2)	2.65	0.55	
Benzhydryl chloride	25	W-E	−122	I	(2)	3.84	0.60	2.19
		W-A	−137	I	(2)	3.79	0.55	2.10
α-Phenethyl chloride	50	W-E	−166	I	(2)	3.92	0.45	1.99
		W-A	−150	I	(2)	3.96	0.49	1.92
Neopentyl bromide	95	(W-E HCOOH)	−162	I	(2)	3.62	0.42	1.96
Pinacolyl-*p*-bromobenzene sulfonate	70	(W-E, W-A Ac_2O-AcOH)	−103	I	(2)	8.05	0.60	2.34
α-Methylallyl chloride	25	(W, W-E)	−161	I	(2)	3.48	0.49	1.99

a) W: water, E: ethanol, M: methanol, A: acetone.
b) The depth of the charge on the halogen atom from the surface.
c) b_* calculated by the present author with Laidler-Landskroener's equation (Eq. 14).

in Method II, L_{AH_2} and L_{AH^-} were evaluated assuming that the charge on every atom in the COOH group is zero. This greatly simplifies the calculation, without affecting the results seriously (compare b's for I and II in Table III). Hence, adipic acid and azelaic acid were treated by Method II. For these two acids calculations were made for the all *trans* form and the di-*cis* or tri-*cis* form (see Appendix II).

The results of the calculations are summarized in Table III. In the last two columns are presented the values of d, the depth of the charge on the oxygen atom in the COO^- group from the surface of the sphere[20]. It is noted that d's are nearly constant for the acids in spite of the wide variation in b's.

S_N1 Solvolysis. — The influence of the dielectric constant of the medium upon the rate constant k_1 of S_N1 solvolysis has been discussed by Shingu and Okamoto[21]. The linear relationship between $\log k_1$ and $1/D$ was interpreted quantitatively on the basis of an ion-pair model for the activated complex. In the present study, the activated complex itself is treated as a sphere, within which positive and negative charges are assumed to be located at the centers of the carbon and halogen atoms, respectively. The equation to be used is as follows;

$$\frac{d \log k_1}{d(1/D)} = \frac{\varepsilon^2}{2.303\,(2kT)}\left(\frac{L_A}{b_A} - \frac{L_*}{b_*}\right) \qquad (16)$$

where L_A/b_A and L_*/b_* are concerned with the reactant molecule and the activated complex, respectively. In the manner similar to that of the preceding example, the radius of the activated complex b_* was determined to fit the experimentally obtained slope.

For *tert*-butyl chloride and *tert*-butyl bromide the charge distributions were calculated from the observed dipole moments[22,23] considering the inductive effect[24], and the L_A's were estimated. Since it was found that they are sufficiently small (see Appendix II), and the final values of b_* are only slightly affected by ignoring them (by 0.01 Å or less), $L_A/\bar{b_A}$ in Eq. 16 was approximated to zero for the present example. In the activated complexes the carbon-halogen bond lengths were assumed to be increased 50% (Model I) or 10% (Model II) over the normal bond lengths. In calculating L, two types of charge distribution were assumed for the activated complexes. Type I

20) d is equal to b minus r for the oxygen atom.
21) H. Shingu and K. Okamoto, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **78**, 558 (1957).
22) L. E. Sutton, "Determination of Organic Structures by Physical Methods", edited by E. A. Braud and F. C. Nachod, Academic Press, New York (1955), p. 392.
23) C. P. Smyth, "Dielectric Behavior and Structure", McGraw-Hill, New York (1955).
24) R. P. Smith, T. Ree, J. L. Magee and H. Eyring, *J. Am. Chem. Soc.*, **73**, 2263 (1951); R. P. Smith and H. Eyring, ibid., **74**, 229 (1952).

consists of a unit negative charge on the halogen atom, and compensating positive charges distributed on the remaining atoms in the same proportion to those in the normal molecules. In Type 2, simplification was made, so that a unit positive charge is localized solely on the α-carbon atom, neglecting the charges on the remaining atoms except for a unit negative charge on the halogen atom. Calculation showed that the solved b_*'s are somewhat smaller in Type 2. Since the difference in b_*'s is not serious, the other five compounds were treated with the simplified models (Type 2), assuming that the carbon-halogen bond lengths in the activated complexes are increased 50% over the normal ones (Model I).

The results are listed in Table IV, where b_*'s calculated with Laidler-Landskroener's equation (Eq. 14) are also included. d indicates the depth of the charge on the halogen atom from the surface of the sphere. It is noticed that d's, in contrast to b_*'s, are not affected appreciably by the choice between Models I and II.

III. Acid and Base Catalyzed Hydrolyses of Esters, Amides and Anilides.—As an example of bimolecular reactions, and for the purpose of comparing the applicability of the present theory with that of Laidler-Landskroener's, the acid and base hydrolyses of certain esters, amides and anilides are treated along the same line and on exactly the same models for the activated complexes as those they proposed[6]. The relevant equation is

$$\frac{\mathrm{d}\log k}{\mathrm{d}(1/D)} = \frac{\varepsilon^2}{2.303(2kT)}\left(\frac{1}{b_A} - \frac{L_*}{b_*}\right) \qquad (17)$$

where k denotes the bimolecular rate constant, b_A the radius of a hydronium or hydroxyl ion, and L_* and b_* refer to the activated complex. In accordance with Laidler-Landskroener's treatment, it was assumed that the hydronium and the hydroxyl ions are unipolar ions (the charge is located at the center of the ion) with the radii of 1.70 Å[25] and 1.40 Å[26],

TABLE V. THE INFLUENCE OF SOLVENT UPON THE RATE CONSTANTS OF THE ACID AND BASE HYDROLYSES OF ESTERS, AMIDES AND ANILIDES

Compound	Catalyst	Temp. °C	Solvent[a]	$\frac{\mathrm{d}\log k}{\mathrm{d}(1/D)}$	b_*(Å) L-L theory[b]	b_*(Å) Present theory C[c]	b_*(Å) Present theory P[d]
Ethyl acetate	H_3O^+	35	W-D	−11	4.6	5.39	4.29
		35	W-A	−34	4.1	5.08	3.97
Methyl propionate	H_3O^+	35.3	W-A	−31	4.2	5.11	4.00
Ethyl acetate	OH^-	25	W-D	−24	1.9	4.47	3.56
			W-A	−32	1.8	4.45	3.51
			W-*i*Pr	−50	1.7	4.41	3.42
			W-*t*Bu	−61	1.7	4.39	3.38
			W-*n*Pr	−74	1.6	4.37	3.34
			W-E	−96	1.5	4.35	3.29
			W-M	−155	1.3	4.31	3.20
Formanilide	H_3O^+	30	W-E	−13.8	3.7	4.33	3.65
Form-*o*-toluidide	H_3O^+	30	W-E	−14.5	3.7	4.32	3.64
Form-*p*-toluidide	H_3O^+	40	W-E	−13.8	3.7	4.32	3.65
o-Nitro-acetanilide	H_3O^+	80	W-E	− 5.5	3.9	4.42	3.76
Acetanilide	H_3O^+	80	W-E	− 5.4	3.9	4.42	3.76
Aceto-*m*-toluidide	H_3O^+	80	W-E	− 7.8	3.8	4.38	3.72
Acetamide	H_3O^+	80	W-E	−42.1	3.2	4.09	3.37
Propionamide	H_3O^+	80	W-E	−39.7	3.2	4.10	3.38
Acetamide	OH^-	80	W-E	−71.3	1.6	4.39	3.34
Propionamide	OH^-	80	W-E	−68.4	1.6	4.39	3.35

a) W: water, E: ethanol, M: methanol, D: dioxane, *i*Pr: isopropanol, *n*Pr: *n*-propanol, *t*Bu: *tert*-butanol.
b) Values calculated by Laidler and Landskroener[6].
c) Center: Carbonyl carbon atom.
d) Center: Point P depicted in Fig. A-5 in Appendix II.

25) J. D. Bernal and R. H. Fowler, *J. Chem. Phys.*, **1**, 515 (1933).
26) J. Buchanan and S. D. Hamann, *Trans. Faraday Soc.*, **49**, 1425 (1953).

respectively. Hence their L's are unity. L's of the substrate molecules were also approximated to zero. By Eq. 17, b_*'s were determined to fit the observed slopes using their models for the activated complexes (see Fig. A-5 in Appendix II). The results are listed in Table V. In the column of b_*, L-L indicates Laidler-Landskroener's values[6], and C or P denotes the results of the present study when the center of the activated complex is taken at the carbonyl carbon atom as in Laidler-Landskroener's calculation, or at the point P depicted in the models, respectively.

Discussion

In the previous theories of the solvent effect[1-5], an ion or a molecule was treated as a uniformly charged sphere or a dipole, respectively, disregarding the actual distribution of charges within them. It is more reasonable to treat the reactants and the activated complex as those having the particular distribution of charges. In this sense, Laidler-Landskroener's treatment[6] and the present one may be said to be more accurate and general than the previous treatments.

However, as seen from Tables IV and V, the radii of the activated complexes calculated according to Laidler-Landskroener's formula are unreasonably too small for the S_N1 solvolysis and the base hydrolysis of esters and amides. These values of radii would lead to an improbable situation where some of the charges are extruded out of the sphere, which is obviously inconsistent with the Kirkwood's model on which the theory is based. On the other hand, the application of the present theory to the above three examples has revealed that the radii of the species concerned are of quite reasonable magnitude as are to be expected from the structures of the species. It is evident from Example III that this discrepancy has come from the approximation made in the derivation of Laidler-Landskroener's equation in which the higher terms were dropped in the sum in Eq. 1, the other circumstances being equal in both treatmnets. The difference between the two theories may be most clearly illustrated in the following example. Suppose a unimolecular reaction in which the reactant is a neutral molecule having no distribution of charge and the activated complex is such as is shown in Fig. 2. Since both the net charge and the dipole moment are zero, the slope $\mathrm{d}\log k/\mathrm{d}(1/D)$ predicted by Laidler-Lands-kroener's equation (Eq. 14) is zero. On the other hand, the slope predicted by the present theory at 25°C is

$$\frac{\mathrm{d}\log k}{\mathrm{d}(1/D)} = -\frac{\varepsilon^2}{2.303(2kT)}\left(\frac{L}{b}\right) = -121.7 \times \frac{11.8}{5}$$

$$= -286 \qquad (18)$$

The value of L depends sensitively upon the depth of the charge, d, at which the charge is located within the sphere. It is worthwhile to examine the values of d determined to fit the observed slopes in the above examples, since they serve as a more rigorous check for the validity of the theory than the radii do.

In treating the interaction energies between charges on organic ions, as reflected in pK's of dibasic acids and amino acids, based on the Kirkwood's model, Tanford[27] showed that the interaction energies critically depend upon the depth, d, of the charges, and it is necessary that d be assigned a value close to 1.0 Å to account for experimentally observed interaction energies. He also found[13] that in calculating titration curves for simple protein model molecules using the Kirkwood model, merely by placing charges about 1 Å below the surfaces the titration curves can be obtained with a slope whose order of magnitude is in the best agreement with that observed experimentally.

Application of the present theory to the dissociation of carboxylic acids showed, as seen from Table III, that d's of the charge on the oxygen atom in the carboxylate group are of nearly constant value around 1.0 Å within the error of the slope[28]. It is of interest to see that these values of d are in good agreement with Tanford's value obtained from the calculation of interaction energies stated above, although different lines of approach have been made. These values of d are comparable to the van der Waals radius (1.40 Å) or the ionic radius (1.4 Å) of oxygen atom[29]. Considering the simplicity of the Kirkwood model, the agreement may be considered to be satisfactory.

The depths, d's calculated for the S_N1 solvolysis are also nearly constant for the series of compounds, but are considerably smaller compared with the van der Waals or ionic radii of the halogen atoms[29]. However, we have good reasons to accept the first example (variation of pK of carboxylic acids with D) as the most reliable one, for in the remaining two examples the structures and the charge configurations of the activated complexes are essentially speculative in nature and are not so definitely established. In addition there is

27) C. Tanford, *J. Am. Chem. Soc.*, **79**, 5348 (1957).
28) The error in the slope may amount to ca. 10%, which leads to an uncertainty of 0.2 Å in b or d.
29) L. Pauling, "The Nature of the Chemical Bond", Cornell University Press, New York (1940),

some possibility that an influence other than that of the dielectric constant might be included in the experimentally determined slopes. In fact, for the S_N1 solvolysis of *tert*-butyl chloride, Okamoto and Shingu[30] have pointed out that the solvent molecules containing one or more hydroxyl groups have some specific effect in accelerating the solvolytic reaction, and this effect is to be distinguished from the dielectric constant effect. Therefore, the observed slopes may be, at least in part, attributed to this specific effect caused by changing the composition of mixed solvents. The true slopes may be less negative, which may lead to larger values of d in the above calculations.

Since the present theory has been developed for zero ionic strength, a comment will be needed about the influence of ionic strength, μ, on the theoretical expressions. According to Kirkwood, the work of charging W of the spherical molecule stated above in the presence of external ions in the solvent is[7]

$$W = W_0 + W_{\text{ion}} \tag{19}$$

where W_0 is a work of charging at zero ionic strength which was already expressed in Eq. 1, and W_{ion} is a term arising from the presence of external ions and is given in the following formula[7,12]

$$W_{\text{ion}} = -\frac{\varepsilon^2}{2a}\sum_{k=1}^{M}\sum_{l=1}^{M}\xi_k\xi_l C_{kl} \tag{20}$$

where a is the mean distance of closest approach of ions in the solvent to the solute species and C_{kl} is a complex function of D, D_t, a, μ and the charge configuration of the species, and is difficult to manage in a simple manner. However, at low ionic strength[31] it is a good approximation to write C_{kl} as

$$C_{kl} \simeq \frac{1}{D}\cdot\frac{\kappa a}{1+\kappa a} \tag{21}$$

where κ is the Debye-Hückel parameter

$$\kappa = \sqrt{8\pi N\varepsilon^2\mu/1000DkT} \tag{22}$$

and is also a function of D. In a manner similar to that of Eq. 7 through Eq. 13, an expression for $\mathrm{d}\log k/\mathrm{d}(1/D)$ in the presence of ions, corresponding to, e. g., Eq. 12, is obtained as follows;

$$\frac{\mathrm{d}\log k}{\mathrm{d}(1/D)} = \frac{\varepsilon^2}{2.303(2kT)}\left[\left(\frac{L_A}{b_A}+\frac{L_B}{b_B}-\frac{L_*}{b_*}\right) - \frac{\mathrm{d}}{\mathrm{d}(1/D)}\left(\frac{\Phi_A}{a_A}+\frac{\Phi_B}{a_B}-\frac{\Phi_*}{a_*}\right)\right] \tag{23}$$

30) K. Okamoto and H. Shingu, reported at the Symposium on Mechanisms of Organic Reactions, Tokyo, 1959.

31) Ionic strength up to κa 0.5 with charges near the surface, and to even higher values of κa for charges appreciably below the surface[12].

where

$$\Phi = \sum_{k=1}^{M}\sum_{l=1}^{M}\xi_k\xi_l C_{kl} \simeq \frac{Z^2}{D}\cdot\frac{\kappa a}{1+\kappa a} \tag{24}$$

since $\sum_{k=1}^{M}\sum_{l=1}^{M}\xi_k\xi_l$ is equal to the square of the total net charge Z^2 of the species. The second term in the right hand side in Eq. 23 represents a contribution of the term dependent on ionic strength to the slope. The contribution at various ionic strengths was estimated for the reaction

$$A + B \rightleftharpoons X_* \rightarrow \text{Products}$$

for which $a_A = a_B = 5$ Å and $a_* = 6$ Å, at 25°C and $D = 78.5$. The results are summarized in Table VI, where φ designates the contribution of the term dependent on ionic strength to the slope.

$$\varphi \equiv \frac{\varepsilon^2}{2.303(2kT)}\cdot\frac{\mathrm{d}}{\mathrm{d}(1/D)}\left(\frac{\Phi_A}{a_A}+\frac{\Phi_B}{a_B}-\frac{\Phi_*}{a_*}\right) \tag{25}$$

TABLE VI. THE CONTRIBUTION OF THE TERM DEPENDENT ON IONIC STRENGTH TO THE SLOPE

	φ		
μ	$Z_A=+1$ $Z_B=+1$	$Z_A=+1$ $Z_B=-1$	$Z_A=\pm1$ $Z_B=0$
0.01	− 9.2	9.9	0.17
0.04	−14.7	16.7	0.49
0.09	−18.2	21.6	0.83
0.25	−23.1	28.1	1.26
0.49	−24.9	32.2	1.83
1.0	−26.0	37.0	2.75

φ is seen to be positive or negative according as the charges of the reactant ions are of opposite or like sign, and to be small if one of the reactants has no net charge. Although the φ is only approximate especially at high ionic strengths, it may serve as a rough estimate of the contribution of the term dependent on ionic strength to the slope.

Summary

A theory has been developed on the relationship between the equilibrium constant or the rate constant of a reaction and the dielectric constant of the solvent based on the Kirkwood's model. It is applicable to equilibrium states or reaction processes in which ions or molecules having arbitrary charge distributions are involved, including enzyme reactions. The theory predicts that there is a linear relationship between the logarithm of the equilibrium or rate constant and the reciprocal of the dielectric constant of the solvent. The slope is determined by the new parameter L, termed the charge

configuration function, and the radius of each species involved. L of a species is a function of the magnitudes and the relative positions of charges in it and its radius. Tables have been presented in order to facilitate the evaluation of L.

The applicability of the theory has been tested for the dissociation of carboxylic acids, the S_N1 solvolysis of organic halides, and the acid and base hydrolyses of esters, amides and anilides, and the results are found to be satisfactory.

The influence of ionic strength on the theoretical expressions has also been discussed.

The author is deeply grateful to Professor Sōzaburo Ono for his kind guidance and discussion throughout this work. He is also indebted to Dr. Kunio Okamoto of Kyoto University and to Dr. Ichiro Moritani and Dr. Tsutomu Ishimura of Osaka University for their helpful discussions. He thanks Professor Wasaburo Jōno of Kyoto University for his interest and encouragement.

Laboratory of Biophysical Chemistry
College of Agriculture
University of Osaka Prefecture
Sakai, Osaka

Appendix I

Evaluation of L. — The charge configuration function L of the hypothetical dipolar ion depicted in Fig. 2 is evaluated as follows; L is written as

$$L=\xi_1^2f_{11}+\xi_2^2f_{22}+\xi_3^2f_{33}+\xi_4^2f_{44}+2\xi_1\xi_2g_{12}+2\xi_1\xi_3g_{13}+2\xi_1\xi_4g_{14}+2\xi_2\xi_3g_{23}+2\xi_2\xi_4g_{24}+2\xi_3\xi_4g_{34}$$

where

$$\xi_1=\xi_3=-\xi_2=-\xi_4=1$$

And

$$b=5\ \text{Å}$$

$$r_1=r_2=r_3=r_4=4\ \text{Å}$$

$$x_{11}=x_{22}=\cdots\cdots=x_{12}=\cdots\cdots=x_{34}=(r/b)^2=0.64$$

$$\alpha_{12}=\alpha_{14}=\alpha_{23}=\alpha_{34}=\cos 90^\circ=0$$

$$\alpha_{13}=\alpha_{24}=\cos 180^\circ=-1$$

The values of $f(x)$ and $g(x,\alpha)$ are given from Tables I and II as follows;

$$f_{11}=f_{22}=f_{33}=f_{44}=3.98$$

$$g_{12}=g_{14}=g_{23}=g_{34}=0.743$$

$$g_{13}=g_{24}=0.447$$

Hence L is evaluated to be

$$L=4\times3.98-2\times4\times0.743+2\times2\times0.447=11.76$$

Appendix II

The geometry and the charge distribution necessary for the calculation of L's in the above three examples are summarized here. r's and θ's are readily computed from them with trigonometry. A–B designates the distance between atoms A and B, ξ_A the valence of the charge on A. The positions of the centers of the species are also shown. The procedure of determining b to fit the observed slope is illustrated for the acetate ion.

Dissociation of Carboxylic Acids.

1) H_2O :— O–H=0.96 Å, ∠HOH=105°[29]. $\xi_0=-0.66$, $\xi_H=0.33$ (calculated from the dipole moment[29]). Center: Oxygen atom. $b=1.40$ Å[32]. L calculated = 0.263.

2) H_3O^+ :— O–H=1.01 Å, ∠HOH=117°[33] (pyramidal structure with equivalent H). $\xi_H=0.50$, $\xi_0=-0.50$. Center: Oxygen atom. $b=1.70$ Å. L calculated = 1.266.

3) Carboxyl group:—The structure of the carboxyl group[23] for acetic acid, benzoic acid and succinic acid (Method I) is shown in Fig. A-1. All the atoms are assumed to lie on the same plane[34]. The charge distribution is shown in the individual articles below.

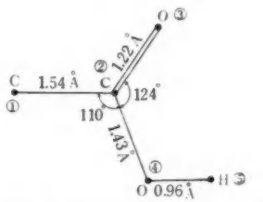


Fig. A-1. The structure of the carboxyl group. The atoms are numbered as indicated.

4) Carboxylate group:—The structure[35] and the charge distribution of the carboxylate group is depicted in Fig. A-2. This was used as the carboxylate model in every acid anion throughout this

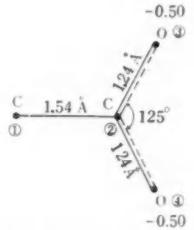


Fig. A-2. The structure of the carboxylate group. The atoms are numbered as indicated. The valences of the charges on the oxygen atoms are also shown.

32) R. W. Gurney, "Ionic Processes in Solution", McGraw-Hill, New York (1953), p. 16.
33) S. Seki et al., "Gendai Kagaku", VIG, Iwanami, Tokyo (1956), p. 51.
34) R. F. Curl, *J. Chem. Phys.*, **30**, 1529 (1959).
35) G. A. Jeffrey and G. S. Parry, *J. Chem. Soc.*, **1952**, 4864.

example. The two oxygen atoms are equivalent[36]. It was assumed that the four atoms are coplanar, and $\xi_1=\xi_2=0$, $\xi_3=\xi_4=-0.50$.

5) Acetic acid (AH):—The charge distribution was calculated from the bond moments[23] and the dipole moment[23] considering the inductive effect[24]. $\xi_1=0.110$, $\xi_2=0.409$, $\xi_3=-0.412$, $\xi_4=-0.435$, $\xi_5=0.325$. Center: Carboxyl C atom. $b=3.00$ Å. L calculated $=0.097$.

6) Acetate ion (A^-): —Center: Point on the C_1-C_2 line at 0.64 Å apart from C_2 atom. L_{A^-}/b_{A^-} calculated for various values of b_{A^-} are as follows;

b_{A^-}(Å)	2.40	2.70	3.00	3.30
L_{A^-}/b_{A^-}(Å^{-1})	0.689	0.529	0.437	0.376

Insertion of the observed slope $d(pK)/d(1/D)$, L's and b's for H_2O, H_3O^+ and HA into Eq. 15 gives

$$\frac{L_{A^-}}{b_{A^-}}=\frac{118}{121.7}+\frac{0.263}{1.40}+\frac{0.097}{3.00}-\frac{1.266}{1.70}=0.445\ (\text{Å}^{-1}) \qquad \text{(A-1)}$$

L_{A^-}/b_{A^-} is plotted against b_{A^-}, and the value of b_{A^-} which fulfils Eq. A-1 can be obtained graphically. b_{A^-} thus obtained is 2.97 Å.

7) Benzoic acid (AH):—The charge distribution was determined to fit the observed dipole moment[23] assuming that ξ_3, ξ_4 and ξ_5 are the same as those in acetic acid and the remaining charges are shared by the C_2 atom and the center of the ring, Q. $\xi_2=0.455$, $\xi_3=-0.412$, $\xi_4=-0.435$, $\xi_5=0.328$, $\xi_Q=0.064$. Center: Point on the C_2-Q line at 1.88 Å apart from C_2. $b=4.60$ Å. L calculated $=0.102$.

8) Benzoate ion (A^-): —Center: Point on the C_2-Q line at 2.32 Å apart from C_2 atom.

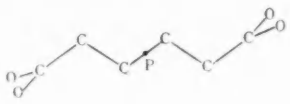
9) Succinic acid (AH_2) (Method I):—*trans* Form, symmetric with respect to the middle point of the C-C bond in CH_2-CH_2 group, was assumed. The four carbon atoms lie on the same plane, to which the planes including each carboxyl group are perpendicular. $\xi_1=0$, $\xi_2=0.519$, $\xi_3=-0.412$, $\xi_4=-0.435$, $\xi_5=0.328$. The charge distribution in the two carboxyl groups are the same. Center: Center of symmetry. $b=4.60$ Å. L calculated $=0.210$.

10) Bisuccinate ion (AH^-) (Method I):—*trans* Form similar to AH_2 except that one of the carboxyl groups is replaced by carboxylate group. Center: C atom adjacent to the carboxyl C atom. The charge distribution in the carboxyl group is the same as that in AH_2.

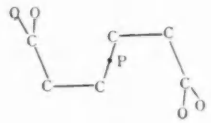
11) Succinate ion ($A^=$) (Method I):—Symmetric *trans* form similar to AH_2 except that the carboxyl groups are replaced by the carboxylate groups. Center: Center of symmetry.

12) Succinic acid (AH_2, AH^-, $A^=$)(Method II): —The geometry of each species is the same as that of the corresponding species treated in Method I. The charges on all the atoms in COOH group were assumed to be zero.

13) Adipic acid (AH_2, AH^-, $A^=$) (Method II): —The structures assumed are schematically depicted in Fig. A-3. In both forms, all the carbon atoms lie on the same plane, to which the two planes including each carboxyl group are perpendicular.

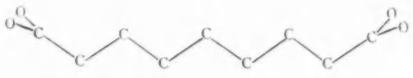


all *trans* form



di-*cis* form

Fig. A-3. Schematic representation of the structures of adipic acid.



all *trans* form

tri-*cis* form

Fig. A-4. Schematic representation of the structures of azelaic acid.

Center: Middle point of C-C bond, P, as shown in Fig. A-3.

14) Azelaic acid (AH_2, AH^-, $A^=$) (Method II): The structures assumed are depicted in Fig. A-4. All the carbon atoms were assumed to lie on the same plane, to which the two COOH planes are perpendicular. Center: Center of the rectangle formed by the four oxygen atoms in $A^=$ form.

SN1 Solvolysis.

In *tert*-butyl chloride and *tert*-butyl bromide, the atoms are numbered as follows; methyl carbon atom 1, α-carbon atom 2, and halogen atom (X) 3. The carbon atoms were all assumed to be tetrahedral. The center was chosen at C_2 atom in both normal state and activated state.

1) *tert*-Butyl chloride.

Normal state: —C-X=1.76 Å, C-C=1.54 Å. $\xi_1=0.0417$, $\xi_2=0.0882$, $\xi_3=-0.213$. (The small charges

36) L. J. Bellamy, "The Infra-red Spectra of Complex Molecules", Methuen & Co., London (1958), p. 161.

on the methyl hydrogen atoms were included in ξ_1). $b=3.00$ Å. L calculated $=0.048$.

Activated state :— C-X $=2.64$ Å (Model I) ; C-X $=1.94$ Å (Model II). $\xi_1=0.196$, $\xi_2=0.414$, $\xi_3=-1.0$ (Type 1), $\xi_1=0$, $\xi_2=1.0$, $\xi_3=-1.0$ (Type 2).

2) *tert*-Butyl bromide.

Normal state :— C-X $=1.91$ Å, C-C $=1.54$ Å. $\xi_1=0.0403$, $\xi_2=0.085$, $\xi_3=-0.206$. $b=3.00$ Å. L calculated $=0.057$.

Activated state :— C-X $=2.86$ Å (Model I) ; C-X $=2.10$ Å (Model II). $\xi_1=0.195$, $\xi_2=0.414$, $\xi_3=-1.0$ (Type 1); $\xi_1=0$, $\xi_2=1.0$, $\xi_3=-1.0$ (Type 2).

The following five compounds were treated by the simplified type of charge distribution for the activated complexes (Type 2), $\xi_2=1$, $\xi_3=-1$, ignoring every small charge on the remaining atoms. The C-X bond distance in the activated complex was taken as 2.86 Å for bromide, and 2.64 Å for chloride (Model I).

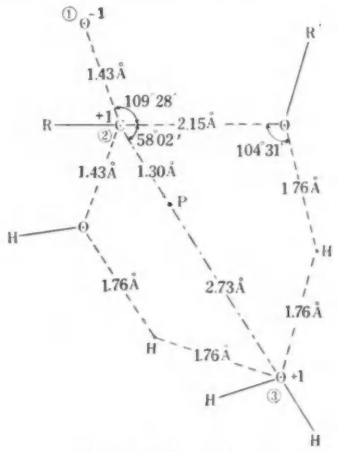
3) Neopentyl bromide :—Center : Middle point between α-C and β-C atoms.

4) Benzhydryl chloride :—Center : Middle point between the two aromatic C atoms to which the α-C atom is attached.

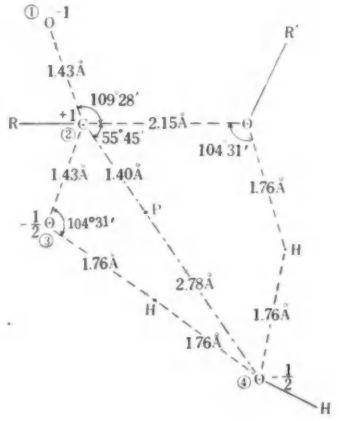
5) α-Phenethyl chloride :—Center : C_1 atom of phenyl group.

6) Pinacolyl-*p*-bromobenzenesulfonate :—Center : Sulfur atom.

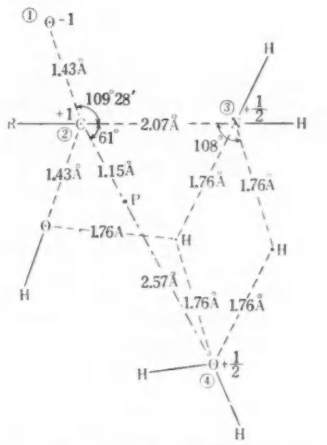
7) α-Methylallyl chloride:—Center : Middle point between the α-C atom and the adjacent double bonded C atom.



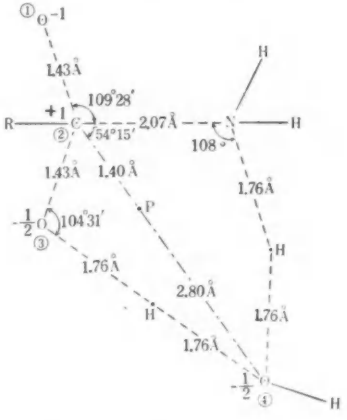
(1) Acid hydrolysis of an ester



(2) Base hydrolysis of an ester



(3) Acid hydrolysis of an amide or anilide



(4) Base hydrolysis of an amide or anilide

Fig. A-5. Laidler-Landskroener's models for the activated complexes in the acid and base hydrolyses of esters, amides and anilides.

III. Acid and Base Hydrolyses of Esters, Amides and Anilides.

Laidler-Landskroener's models for the activated complexes are shown in Fig. A-5. It was assumed that the C_2 atoms are tetrahedral, and the following four atoms lie on the same plane;

O_1, C_2, O_3 and O(esteric) in Model (1)
O_1, C_2, O_4 and O(esteric) in Model (2)
O_1, C_2, N_3 and O_4 in Model (3)
O_1, C_2, N and O_4 in Model (4).

The center was chosen at C_2 atom or at point P depicted in the models (cf. Table V).

*Theory of the Influence of the Dielectric Constant on the Rate of Reaction in Solution with Application to Enzyme Reactions. II. Application of the Theory to Enzyme Reactions**

By Keitaro HIROMI

(Received December 18, 1959)

The enzyme molecule in solution has a number of charged groups, some of which play an essential role in the catalytic reactions. Rates of enzyme reactions, therefore, may be expected to be influenced by the dielectric constant of a medium as are generally observed in reactions involving simple ions or dipolar molecules. From this viewpoint, rates of several enzyme reactions have been measured in aqueous mixed solvents, and the results have been discussed in relation to the dielectric constant of the medium[1-8]. In treating catalase and peroxidase reactions, Ogura et al.[1] and Tonomura[3] considered the enzyme or the enzyme-substrate complex as a dipole, to which the expression for the reaction between dipolar molecules[9] was applied. On the other hand, for α-chymotrypsin[2] and myosin[4] catalyzed reactions, Laidler and his coworkers applied the coulombic law to the interaction between the charges of the substrate and the essential groups of the enzyme. However, since it is supposed that the reaction may occur in a local region near the surface of a large cavity of low dielectric constant formed by the enzyme molecule in solution, these treatments may be said to be over-simplified, and the validity of the conclusions deduced from them may be subject to criticism.

The Kirkwood's model for a spherical solute molecule having an arbitrary charge distribution, on which a general theory of the solvent effect has been developed in the preceding paper[10], seems to be appropriate as a model for a protein in solution. In this paper, the theory will be extended to include the solvent effect upon enzyme reactions.

Theoretical

Enzyme reactions in general consist of two stages, formation of an intermediate enzyme-substrate complex, and its breakdown into products and free enzyme.

$$E + S \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} ES \xrightarrow{k_2} E + P$$

The Michaelis constant K_m and the rate constant of breakdown k_2 can be obtained in the usual manner[11]. In certain cases in which $k_{-1} \gg k_2$, K_m may be identified with the reciprocal of the equilibrium constant K for the formation of ES complex[12]. In some instances, k_1 or k_{-1} can be measured directly[13], or may

* Presented in part at the Symposium on Enzyme Chem., Sapporo, July, 1958.

1) Y. Ogura, A. Hattori, Y. Tonomura and S. Hino, *Symposia on Enzyme Chem. (Kōso Kagaku Shimpojiumu)*, **5**, 28 (1950); Y. Ogura and H. Hasegawa, ibid., **7**, 79 (1952).

2) M. L. Barnard and K. J. Laidler, *J. Am. Chem. Soc.*, **74**, 6099 (1952).

3) Y. Tonomura, *J. Japan Biochem. Soc.*, (*Seikagaku*), **25**, 175 (1953).

4) K. J. Laidler and M. C. Ethier, *Arch. Biochem. Biophys.*, **44**, 338 (1953).

5) R. Lumry and E. L. Smith, *Discussions Faraday Soc.*, **20**, 105 (1955).

6) S. Ono, K. Hiromi and Y. Sano, Presented as a lecture at the 8th Annual Meeting of Chem. Soc. Japan, Tokyo (1955); S. Ono, K. Hiromi and Y. Sano, to be published.

7) A. Stockell and E. L. Smith, *J. Biol. Chem.*, **227**, 1 (1957).

8) T. Hosoya, *J. Biochem.* (Tokyo) in press.

9) S. Glasstone, K. J. Laidler and H. Eyring, "The Theory of Rate Processes", McGraw-Hill New York (1941), p. 419.

10) K. Hiromi, This Bulletin, **33**, 1251 (1960).

11) For example, H. Lineweaver and D. Burk, *J. Am. Chem. Soc.*, **56**, 658 (1934).

12) M. F. Morales, ibid., **77**, 4169 (1955); K. J. Laidler, *Trans. Faraday Soc.*, **51**, 540 (1955).

13) B. Chance, "Investigation of Rates and Mechanisms of Reactions", in "Technique of Organic Chemistry", edited by A. Weissberger, Interscience Publishers, New York (1953), Vol. VIII, p. 627.

be obtained from the analysis of overall reaction rates[14]. These kinetic parameters, k's and K, may be influenced by the dielectric constant D of the solvent, and equations similar to Eqs. 12 and 13 in the preceding paper[10] may be available. They involve the charge configuration functions L's of the substrate, enzyme, ES complex and/or ES_*, the activated complex of ES. The L's of the last three species involve the contribution of all the charges of the enzyme molecule, which we cannot evaluate explicitly because of our lack of knowledge about their spatial distribution over the enzyme molecule. However, as will be seen shortly, the following treatment may be allowable. Instead of L, the complete charge configuration function involving all the charges in the enzyme, we may use L^0, the effective charge configuration function concerned only with the small number of charges of special interest belonging to a limited reaction region, disregarding all the remaining charges in the enzyme which are considered to be kept fixed and not to come into play during the reaction process.

Consider the equilibrium, for example,

$$E + S \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} ES$$

The variation of the equilibrium constant $K(=k_1/k_{-1})$ with D can be written as[10]

$$\frac{d \log K}{d(1/D)} = \frac{\varepsilon^2}{2.303(2kT)}\left(\frac{L_S}{b_S} + \frac{L_E}{b_E} - \frac{L_{ES}}{b_{ES}}\right) \tag{1}$$

where b's are the radii of the substrate, enzyme and ES complex, as indicated by the suffixes, L's are the corresponding complete charge configuration functions which contain the contribution of all the charges in E and ES, and the other symbols have the same meaning as those in the preceding paper[10]. For small substrates b_E and b_{ES} may be equated. Let E and ES have in the total M and M' charges, respectively, and let n of them be kept fixed in position and free from any change during the process. Thus

$$\left.\begin{aligned} M &= n + m \\ M' &= n + m' \end{aligned}\right\} \tag{2}$$

where m and m' are the number of charges of special interest which belong to the local reaction region in E and ES, respectively. L_E and L_{ES} are written as follows[10];

$$\begin{aligned} L_E &= \sum_{k=1}^{M} \xi_k^2 f_{kk} + \sum_{k=1}^{M} \sum_{l \neq k} \xi_k \xi_l g_{kl} \\ &= \sum_{i=1}^{m} \xi_i^2 f_{ii} + \sum_{s=1}^{n} \xi_s^2 f_{ss} + \sum_{i=1}^{m} \sum_{s=1}^{n} \xi_i \xi_s g_{is} \\ &\quad + \sum_{i=1}^{m} \sum_{j \neq i} \xi_i \xi_j g_{ij} + \sum_{s=1}^{n} \sum_{t \neq s} \xi_s \xi_t g_{st} \end{aligned} \tag{3}$$

$$\begin{aligned} L_{ES} &= \sum_{k=1}^{M'} \xi_k^2 f_{kk} + \sum_{k=1}^{M'} \sum_{l \neq k} \xi_k \xi_l g_{kl} \\ &= \sum_{i=1}^{m'} \xi_i^2 f_{ii} + \sum_{s=1}^{n} \xi_s^2 f_{ss} + \sum_{i=1}^{m'} \sum_{s=1}^{n} \xi_i \xi_s g_{is} \\ &\quad + \sum_{i=1}^{m'} \sum_{j \neq i} \xi_i \xi_j g_{ij} + \sum_{s=1}^{n} \sum_{t \neq s} \xi_s \xi_t g_{st} \end{aligned} \tag{4}$$

where the suffixes i and j refer to the charge number of the charges belonging to the m or m' group, s and t the n group, and f and g are the functions of the positions of charges as were defined by Eq. 4 in the preceding paper[10].

Hence we have

$$\begin{aligned} L_E - L_{ES} &= \left(\sum_{i=1}^{m} \xi_i^2 f_{ii} + \sum_{i=1}^{m} \sum_{j \neq i} \xi_i \xi_j g_{ij}\right)_E \\ &\quad - \left(\sum_{i=1}^{m'} \xi_i^2 f_{ii} + \sum_{i=1}^{m'} \sum_{j \neq i} \xi_i \xi_j g_{ij}\right)_{ES} \\ &\quad + \left(\sum_{i=1}^{m} \sum_{s=1}^{n} \xi_i \xi_s g_{is}\right)_E - \left(\sum_{i=1}^{m'} \sum_{s=1}^{n} \xi_i \xi_s g_{is}\right)_{ES} \end{aligned} \tag{5}$$

The last two terms of the right hand side in Eq. 5 are those from the pair-wise contribution between the m or m' charges of special interest and the n remaining fixed charges in E and ES. If the n charges consist of approximately equal numbers of positive and negative charges, and they are distributed at random over the enzyme molecule, these terms may be appreciably small and may be dropped without serious error. Hence

$$\left.\begin{aligned} L_E - L_{ES} &= L^0_E - L^0_{ES} \\ L^0_E &= \left(\sum_{i=1}^{m} \xi_i^2 f_{ii} + \sum_{i=1}^{m} \sum_{j \neq i} \xi_i \xi_j g_{ij}\right)_E \\ L^0_{ES} &= \left(\sum_{i=1}^{m'} \xi_i^2 f_{ii} + \sum_{i=1}^{m'} \sum_{j \neq i} \xi_i \xi_j g_{ij}\right)_{ES} \end{aligned}\right\} \tag{6}$$

where L^0_E and L^0_{ES} are the effective charge configuration functions of E and ES, concerned only with the charges of special interest. Eq. 1 becomes, therefore,

$$\begin{aligned} \frac{d \log K}{d(1/D)} &= \frac{\varepsilon^2}{2.303(2kT)} \\ &\quad \times \left[\frac{L_S}{b_S} + \frac{1}{b_E}(L^0_E - L^0_{ES})\right] \end{aligned} \tag{7}$$

Similarly for the breakdown process of ES, we have

$$\frac{d \log k_2}{d(1/D)} = \frac{\varepsilon^2}{2.303(2kT)}\left(\frac{L^0_{ES}}{b_{ES}} - \frac{L^0_{ES*}}{b_{ES*}}\right) \tag{8}$$

14) H. Gutfreund, *Discussions Faraday Soc.*, **20**, 167 (1955); E. C. Slater, ibid., **20**, 231 (1955).

where L^0_{ES*} is the effective charge configuration function of ES_*. Similar equations for k_1 or k_{-1} if necessary, may readily be obtained.

Application to Some Typical Models

In Table I are summarized the data of the solvent effect upon the kinetic parameter $\mathfrak{K}$ which have been obtained for various enzyme reactions[1-8]. $\mathfrak{K}$ represents the rate constant or equilibrium constant specified in each item. The slopes $d \log \mathfrak{K}/d(1/D)$ have been calculated under the assumption that the observed changes in $\mathfrak{K}$ are due entirely to the variation in dielectric constant of the mixed solvent. The radii of the enzymes are computed from the molecular weight and the specific volume[15] assuming the spherical shape of the enzymes. In many cases, the slopes have considerably large negative values of the order of -200.

It is worthwhile to see whether the present theory can reasonably account for the slopes of the observed order of magnitude, considering only the electrostatic rearrangement in the local reaction region.

For the present, let us confine our attention to the breakdown process of ES complex in which the bond X–Y of a substrate molecule is to be split. During the activation process a change in charge state may occur in the local reaction region which involves the bound substrate molecule and the catalytic center of the enzyme. We may assume that ES and ES_* are spheres of equal radii b_{ES}, and that the substrate molecule is situated near the surface of the sphere. It is reasonable to suppose that in the activated complex the bond to be split is polarized so that a positive charge and a negative charge will be produced on the atoms X and Y forming the bond, as a preliminary stage of bond splitting[9]. It may lead to a change in L^0 by which the slope $d \log k_2/d(1/D)$ will be determined. The increment in L^0, and hence the slope, may depend on the magnitude of the charges produced, their position in the sphere (or the depth of the charges from the surface), and the situation of the neighboring charges. Two typical models for ES_* are shown in Fig. 1. e_3 and e_4 are the charges produced on the atoms X and Y in ES_* at a depth d from the surface of the sphere. e_1 and e_2 are the neighboring charges which represent, for example, the charges of the active site of the enzyme, or the other charges the substrate may bear. In both models, the distance between e_3 and e_4 is taken as 2 Å, which is nearly equal to 50% stretched bond length of C–O or C–N

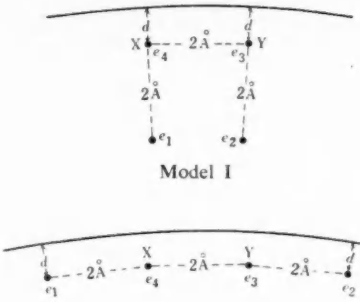


Fig. 1. Models of ES_*.
e_1 and e_2 represent the neighboring charges, and e_3 and e_4 the charges produced on the atoms X and Y in the activated state.

bond over their normal ones. In Model I, e_1 and e_2 are assumed to be situated by 2 Å inside on the lines drawn from the center to the charges e_4 and e_3, respectively. In Model II, e_1 and e_2 are situated on both sides of e_4 and e_3, 2 Å apart from them, respectively, at a depth d from the surface. ξ's denote the valences of the charges. The increment in the effective charge configuration function, ΔL^0_* $(=L^0_{ES*} - L^0_{ES})$, accompanied with the activation of ES complex, can be calculated from Eq. 6, using Tables I and II in the preceding paper[10], and the slope $d \log k_2/d(1/D)$ from Eq. 8. In the present case, calculations were made under the following conditions; i) the charges e_3 and e_4 are both zero in ES, ii) $-\xi_3=\xi_4=1$ or $1/2$ in ES_*, iii) e_1 and e_2 are either positive or negative unit charges and are fixed during the activation process, iv) d is taken as 0.5 Å or 1.0 Å, and v) $b_{ES}=b_{ES*}=21$ Å and the temperature is 25°C. The results are summarized in Table II. It is seen that the slope is determined primarily by the magnitude of the charges produced and by the depth of the charges below the surface, d, and is also dependent upon the situation of the neighboring charges especially in Model II. The slope of the observed order of magnitude is reasonably accounted for, if a pair of charges of opposite sign whose valence is close to unity are produced in the activated complex at the depth of 0.5~1.0 Å below the surface. These values for depth of charges are in reasonable agreement with those determined for the simple systems in the preceding paper[10]. On the other hand, Laidler-Landskroener's equation[16] applied to the above models, yields a slope, at the greatest, of about

15) E. J. Cohn and J. T. Edsall, "Proteins, Amino Acids and Peptides", Reinhold Publishing Co., New York (1943), p. 370.

16) K. J. Laidler and P. A. Landskroener, *Trans. Faraday Soc.*, 52, 200 (1956). (cf. Eq. 14 in the preceding paper).

TABLE I. SOLVENT EFFECTS UPON VARIOUS ENZYME REACTIONS

Substrate	Enzyme	Radius b_E, Å	Temp. °C	k	Solvent[a]	$\frac{d \log k}{d(1/D)}$	Reference No.
Carbobenzoxy-glycyl-L-tryptophan	Carboxy-peptidase	21	25	k_2	W-M	−200	5
					W-E	−180	
					W-D	− 30	
Methyl hydrocinnamate	α-chymo-trypsin	20	25	k_2	W-M	−260	2
				K	W-M	−240	
α-Benzoyl-L-argininamide	Papain	18	38	k_1	W-M	−160	7
			38	k_1	W-*i*Pr	−160	
			5	k_1	W-M	−180	
Soluble starch	Bacterial α-amylase	23	25	k_2	W-M	−210	6
				K	W-M	0	
ATP	Myosin	—	25	k_2	W-M	80	4
				K	W-M	330	
H_2O_2	Catalase	42	0	k_3[b]	W-E	−160	1
H_2O_2+guaiacol	Peroxidase (Japanese radish)	22	25	k[c]	(W-E, W-A W-D)	−200	3
H_2O_2+guaiacol	Peroxidase (Turnip) A_1	23	20	k_4[d]	(W-M, W-A)	−210	8
	A_2	23	20	k_4	W-M	−220	
	D	23	20	k_4	W-M	−250	

a) W: water, M: methanol, E: ethanol, D: dioxane, *i*Pr: *iso*-propanol, A: Acetone.

b) $ES \xrightarrow{s} E + P$ (s: H_2O_2), k_3

c) $ES \xrightarrow{a} E + P$ (a: guaiacol), k

d) The apparent rate constant of the reaction between complex II and guaiacol.

TABLE II. PREDICTION OF THE THEORY FOR TYPICAL MODEL REACTIONS

Charges produced	Model	Neighboring charges ξ_1	Neighboring charges ξ_2	Charge configuration in ES_*[a]	$\Delta L^0_* / d$ = 0.5 Å	$\Delta L^0_* / d$ = 1.0 Å	$-d \log k_2/d(1/D)$ d = 0.5 Å	$-d \log k_2/d(1/D)$ d = 1.0 Å
	I, II	0	0	(○ ●)	46.3	12.8	269	74.2
	I	+1	+1	(○ ● / ○ ○)	46.3	12.8	269	74.2
$\xi_3=-1$	I	+1	−1	(○ ● / ○ ●)	55.1	16.8	320	97.4
$\xi_4=+1$	I	−1	+1	(○ ● / ● ○)	37.5	8.8	218	51.0
	II	+1	+1	(○ ○ ● ○)	46.3	12.8	269	74.2
	II	+1	−1	(○ ○ ● ●)	80.4	34.1	466	198
	II	−1	+1	(● ○ ● ○)	13.0	−8.4	75.4	−48.7
	I, II	0	0	(◦ •)	11.6	3.2	67.3	18.6
	I	+1	+1	(◦ • / ○ ○)	11.6	3.2	67.3	18.6
$\xi_3=-1/2$	I	+1	−1	(◦ • / ○ ●)	16.0	5.2	92.8	30.2
$\xi_4=+1/2$	I	−1	+1	(◦ • / ● ○)	7.2	1.2	41.8	7.0
	II	+1	+1	(○ ◦ • ○)	11.6	3.2	67.3	18.6
	II	+1	−1	(○ ◦ • ●)	28.6	13.8	169	80.0
	II	−1	+1	(● ◦ • ○)	−5.4	−7.4	−31.3	−42.9

a) The charge configurations in ES_* (Models I and II) are schematically illustrated with the white and the black circles. The large white or black circle represents a positive or a negative unit charge, and the small white or black circle a positive or a negative half unit charge, respectively.

-0.8, by a factor of 10^2 less than the observed values.

Discussion

The present theory which is based on the Kirkwood's model may be more suited for treating the solvent effect upon enzyme reactions than the previous ones[1-4], although some simplifications inevitable in dealing with the complex enzyme molecule have still been made. The application of the theory to the typical models has shown that it successfully accounts for the experimentally obtained slopes, considering only the local rearrangement of charges in the enzyme-substrate complex. An advantage of the present theory over the existing theories is that it can deal with any microscopic change in charge state in any position in the spherical solute. It eliminates the disadvantage of the treatment which regards the enzyme or the enzyme-substrate complex as a dipole[1,3], according to which an extraordinarily large dipole moment had to be assigned to the activated complex in order to account for the slope of the observed order of magnitude[17].

At present we shall have to be content with the successful results of the application of the theory to the typical examples stated above. However, when our knowledge has accumulated about the value of d to be assigned in the individual cases, we may reasonably hope that we can have an insight into the microscopic charge configuration in the activated state from the slope obtained by the study of the solvent effect on the reaction rate, which may serve as a tool for elucidating the mechanism of the reaction.

However, in interpreting the experimental data of solvent effect, attention must be paid to the possibility that the observed effects may involve contributions of various factors other than the purely electrostatic effect of dielectric constant on the rate. The factors that may possibly influence the rate by changing the composition of mixed solvents are enumerated as follows; 1) decrease in water concentration which might result in reduction of rates, if the rate determining step involves participation of the water molecules especially for hydrolytic enzymes, 2) shift of the activity-pH curve which may decrease or increase the rate when the experiments are carried out at constant pH[6], 3) possible inhibition by some component of the mixed solvent, 4) irreversible or reversible denatuation of the enzyme caused by the addition of organic solvents, 5) change in viscosity which may influence the rate if the reaction is diffusion controlled[18-20]. In most cases, the addition of organic solvents may lead to a decrease in rates due to these factors, in addition to a decrease in the dielectric constant of the medium. If this is the case, the observed slopes may possibly have been over-estimated in the negative direction, which would lead to larger values of d in the above calculation.

For non-zero values of ionic strength, the slope may be affected by the inclusion of the term dependent on ionic strength as already discussed in the preceding paper[10]. The contribution of this term may be small except for bimolecular processes in which both of the reactants have non-zero net charges. Of course, the contribution is greater at higher ionic strength. A rough estimate reveals that for ionic strength of 0.02 and the radius of the enzyme taken as 20 Å[21], the contribution to the slope may be less than about ± 40 for a bimolecular process between reactants having univalent net charge. It is desirable that the reaction should be carried out at an ionic strength as low as possible, in order to elucidate the effect of dielectric constant on the rate.

The theory may also be applied to the influence of solvent on binding of substrates or inhibitors to the enzyme, although there are restrictions inherent to the Kirkwood's model.

Summary

The theory of solvent effect developed in the preceding paper has been extended to include the enzyme reactions. The application to some typical model reactions which resemble the breakdown of enzyme-substrate complex has shown that the order of magnitude of experimentally observed effects can reasonably be accounted for by the theory.

Factors which may possibly influence rates of enzyme reactions by changing the composition of mixed solvents have also been discussed.

The author expresses his hearty thanks to Professor Sôzaburo Ono for his kind guidance

17) For example, if the slope d log k_2/d$(1/D)$ of the order of -200 is interpreted by the equation for dipolar molecules[1,3,9]

$$\frac{\mathrm{d}\log k_2}{\mathrm{d}[(D-1)/(2D+1)]} = \frac{1}{2.303\,kTb_{ES}^3}(\mu_{ES^+}^2-\mu_{ES}^2)$$

where μ's are the dipole moments, it is necessary that $\mu_{ES^+}^2$ be larger than μ_{ES}^2 by ca. 2.4×10^5 Debye^2, for b_{ES} = 21 Å at 25°C. Accepting this large increase in μ, one might be forcibly led to the conclusion that a profound change in shape or charge distribution in the whole body of the enzyme molecule had been brought about during the activation process.

18) E. Ackerman, G. K. Strother and R. L. Berger, "The Influence of Temperature on Biological Systems", Papers Symposium, Storrs, Conn. (1956), p. 25.

19) R. A. Alberty and G. G. Hammes, *J. Phys. Chem.*, 62, 154 (1958).

20) H. Mitsuda et al., *Symposia on Enzyme Chem.*, (*Kôso Kagaku Shimpojiumu*), in press.

21) Accordingly κa (cf. Eq. 21 in the preceding paper) at room temperature is about unity.

and discussion throughout this work. The author is also indebted to Professor Wasaburo Jôno of Kyoto University for his interest and encouragement.

Laboratory of Biophysical Chemistry
College of Agriculture
University of Osaka Prefecture
Sakai, Osaka

Note: The present theory was recently applied by Hosoya to the interpretation of his data of solvent effect upon the peroxidase reactions. (See Ref. 8.)

Amylose-Iodine Complex. II[1]. *Effects of Inorganic Salts*

By Takashi KUGE and Sôzaburo ONO

(Received March 21, 1960)

The starch-iodine reaction in solution is markedly affected by some conditions (i.e. temperature, history of amylose solution and concentrations of reactants, etc.) under which the reaction takes place. However, the effect of salts on starch-iodine reaction has scarcely been investigated except in the case of studies[2] made at the beginning of this century at which time starch chemistry had not yet been established. The authors have found that certain salts strongly aid the formation of the amylose-iodine complex in a solution containing small amount of iodide ion. Referring these results to the helical theory[3], the authors intend to elucidate the structure of the amylose-iodine complex in solution.

Experimental

Materials.—Amylose was obtained from potato starch by a selective precipitation method with *n*-butyl alcohol and isoamyl alcohol[4]. This amylose was recrystallized twice from boiling water in the presence of excess butanol. All other reagents were of the purest grade commercially available and used without further purification.

Procedure.—Hot water was poured on the recrystallized dry amylose of known amount and the suspension was heated on a water bath, being stirred for one hour in order to complete the dispersion of amylose. After being cooled, the solution was filtered with a glass filter and the filtrate was diluted to a desired volume with water. The above prepared solution, aqueous iodine solution, iodide solution and salt solution were used to make the final solutions which were composed of 0.005~0.015% amylose, 10^{-5}~2×10^{-5} M iodine, 10^{-6}~10^{-5} M potassium iodide and 10^{-5}~10^{-1} M salt. In these compositions the amount of iodine was less than one third of that sufficient to saturate the amylose, and the ratio of iodide to iodine was unusually small, compared with that in a usual case. After being stood for more than one hour, the final colored solutions were used for spectrophotometric measurements.

A nearly saturated solution of iodine which was used to prepare the final solutions was made by shaking iodine crystals in water at room temperature. The concentrations of iodine and iodide ion in this solution were determined immediately before use by measuring their absorbances at wavelengths 288, 350 and 460 mμ[5], since iodide ion is formed by the hydrolysis of iodine.

The spectrophotometric measurements were carried out at room temperature by a Shimadzu type QB-50 spectrophotometer with 1.00 cm. thick silica cells.

Results and Discussion

Iodide ion plays an important role in the amylose-iodine reaction but it remains uncertain whether or not amylose can form an iodide-free iodine complex in solution[6]. To solve this question, an aqueous iodine solution, in which iodic acid, perchloric acid or sulfuric acid was contained to repress any iodide ion due to the hydrolysis of iodine, was added to the amylose solution. Absorption at 350 mμ

1) S. Ono, S. Tsuchihashi and T. Kuge, *J. Am. Chem. Soc.*, **75**, 3601 (1953) is the Part I of this series.
2) For a review, see "Chemistry and Industry of Starch", edited by R. W. Kerr, Academic Press, Inc., New York, (1950), Chapter XVII, p. 455.
3) R. E. Rundle and D. French, *J. Am. Chem. Soc.*, **65**, 554 (1943); R. E. Rundle and R. R. Baldwin, ibid., **65**, 558 (1943); R. E. Rundle, ibid., **69**, 1769 (1957); R. S. Stein and R. E. Rundle, *J. Chem. Phys.*, **16**, 195 (1948).
4) E. J. Wilson, T. J. Schoch and C. S. Hadson, *J. Am. Chem. Soc.*, **65**, 1380 (1943).
5) A. D. Awtrey and R. E. Connick, ibid., **73**, 1842 (1951).
6) E. O. Forster, Ph. D. Thesis, Columbia University, (1951), through J. A. Thoma and D. French, *J. Am. Chem. Soc.*, **80**, 6142 (1958).

due to triiodide ion was not observed in 0.2 M iodic acid solution and the blue coloration due to complex formation did not occur with the amylose solution. On the contrary, when the other acid solutions were used, the absorption due to a small amount of triiodide ion was always observed and the coloration took place. Therefore, the blue amylose-iodine complex seems to be formed only when iodide is in solution.

Gilbert and Marriot[7] have suggested that the constitution of the complex is $(3I_2 \cdot 2I^-)$ or (I_8^{2-}) and this resonating ion is thought to be responsible for the blue color. However, as mentioned in the previous paper[1], the amount of iodide ions necessary for the coloration may be very small compared with that concluded by the above authors. Rundle and French[8] have demonstrated the formation of colored complex by absorption of iodine vapors by helical amylose in solid state[9]. It was, hence, inferred that a large amount of iodide ions is not necessary for the coloration. Recently, Holló and Szejtli[10] have agreed with our opinion concerning the constitution of the complex from amperometric studies. Although there are some evidences[7,11,12] indicating that iodide ions are adsorbed as well as iodine molecules by amylose in solution, it is still in doubt whether they all take part in the resonating polyiodine chains which are responsible for the blue color. The authors are inclined to think that the iodide ions indispensable for the resonating chain are only small fractions of adsorbed iodide ions. The triiodide ion adsorbed on the surface of amylose will be discussed in the following paper[13].

From Fig. 1 it is seen that the amylose-iodine complex has a spectrum having two absorption maxima, each of which is in the visible and ultraviolet region, respectively. The absorption is small in the low concentration of potassium iodide and it increases considerably with increasing iodide concentration until 5×10^{-3} M iodide is reached, after which the blue shift of the spectrum takes place, accompanying the decrease of the visible absorption.

Fig. 2 shows the changes of the spectra with the increase of potassium chloride at constant iodide concentration (10^{-5} M). It is seen that

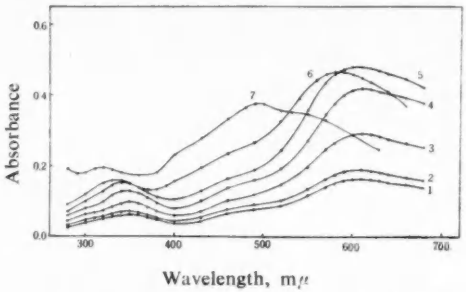


Fig. 1. The absorption spectra of amylose-iodine complexes formed under the various concentrations of potassium iodide: amylose 0.005%, total iodine 1.5×10^{-5} M, potassium iodide (1) 1.3×10^{-5} M, (2) 1.8×10^{-5} M, (3) 6.3×10^{-5} M, (4) 5×10^{-4} M, (5) 5×10^{-3} M, (6) 5×10^{-2} M, (7) 5×10^{-1} M.

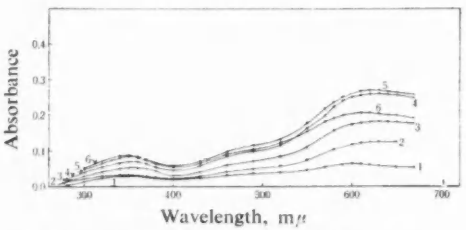


Fig. 2. The absorption spectra of amylose-iodine complexes formed under the various concentrations of potassium chloride: amylose 0.008%, total iodine 1.3×10^{-5} M, potassium iodide 1×10^{-5} M, potassium chloride (1) zero, (2) 5×10^{-4} M, (3) 2.5×10^{-3} M, (4) 1.3×10^{-2} M, (5) 6×10^{-2} M, (6) 1.2×10^{-1} M.

the increase in the concentration of potassium chloride causes the shift of the wavelength of maximum absorption λ_{max}, to a longer wavelength in the visible region, accompanying the considerable increase of absorption.

Certain other salts[14] also aid the complex formation in a manner similar to that of potassium chloride. These results are collected in Fig. 3. Fig. 3 (a) shows the relationship between λ_{max} and the salt concentration and (b) the dependence of absorbance at λ_{max} on the salt concentration, in the visible region. The amylose-iodine complexes formed in the solution, which contains 0.005~0.015% amylose, $10^{-5}\sim2\times10^{-5}$ M iodine and $10^{-6}\sim10^{-3}$ M iodide but no any other salt, has λ_{max} at 600 mμ, and its absorbance E_{max}, is very small. It is seen

7) G. A. Gilbert and J. V. R. Marriott, *Trans. Faraday Soc.*, **44**, 84 (1948).
8) R. E. Rundle and D. French, *J. Am. Chem. Soc.*, **65**, 1707 (1943).
9) Even with dry amylose, traces of iodide ions appeared to be produced by reduction of iodine by impurities (remaining water etc.).
10) J. Holló and J. Szejtli, *Die Stärke*, **10**, 248 (1958).
11) R. S. Higginbotham, *Shirley Institute Memoris*, **23**, 159 (1949).
12) D. L. Molud, *Biochem. J.*, **58**, 593 (1954).
13) T. Kuge and S. Ono, This Bulletin, **33**, 1273 (1960).
14) Salts, whose aqueous solutions show alkaline pH, for example, sodium carbonate, potassium carbonate and disodium hydrogen phosphate, etc., inhibited the coloration. It can be ascribed to the reduction of iodine by hydroxyl ions.

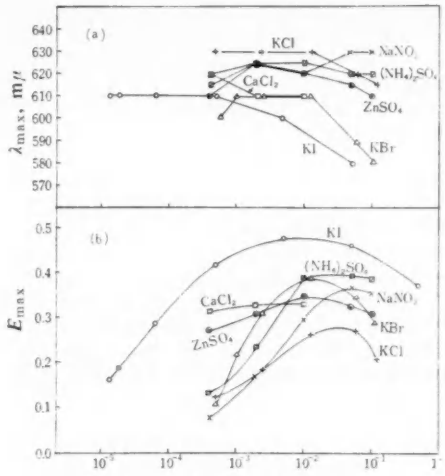


Fig. 3. The effects of various salts on the formation of amylose-iodine complex.

that E_{max} shows a maximum value at 5×10^{-3} M iodide. The maximum value seems to be determined by the iodine concentration in the solution, while the maximum values caused by the additions of the other salts seem to be mainly determined by the iodide concentration. The addition of more iodide than 10^{-2} M causes the large shift of λ_{max} to a shorter wavelength as described in the previous paper[1]. The blue shifts due to additions of an excess of the other salts are small and λ_{max} are always more than 600 mμ. Only bromide is exceptionally similar to iodide in respect of the blue shift. This seems to be caused by the formation of the iodine-bromide ion[15] that would enter the polyiodine chain and shorten the resonating length as iodide ions do[1], but it may not be the case with other ions.

As described above, the complex formation seems to be influenced to a great extent by iodide and the other salts. Iodide, however, has a different role in the complex formation from the other salts; that is, iodide is essential for the reaction, while the other salts aid strongly the formation of the complex from amylose, iodine and iodide ion when the concentration of iodide is as low as 10^{-5} M. Since the decrease of E_{max} observed with the addition of salts over 0.01 M seems to be partly owing to the precipitation of the complexes, this phenomenon will not be further discussed.

15) A marked change of the absorption spectrum of aqueous iodine solution was observed by the addition of bromide or chloride, indicating the strong interaction between iodine and bromide or chloride ion. The other salts hardly altered the spectrum.

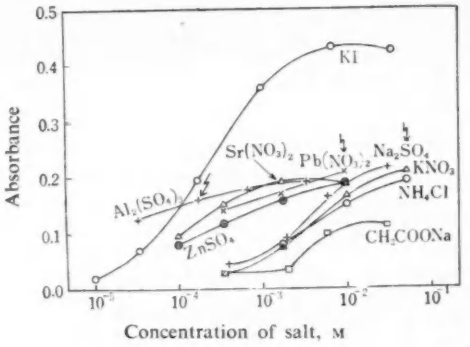


Fig. 4. The changes of absorbance at 610 mμ with increasing concentrations of various salts: amylose 0.007%, total iodine 1.25×10^{-5} M, potassium iodide 1×10^{-5} M.

Fig. 4 shows another series of experiments carried out. It can be seen that the investigated salts are divided into three groups according to their valency. The higher the valency which the cation has, the lower becomes the concentration of the salt efficient for the completion of the amylose-iodine reaction. The arrows in Fig. 4 indicate the concentrations at which the precipitant of the complex becomes apparent to the naked eye. This concentration also changes in a regular fashion with the valency of cation. These facts may be connected with the Schulze-Hardy rule. Thus it seems quite natural to correlate the increase in absorbance of the complex by added salts with the increase of the aggregation of amylose through the neutralization of the charge on amylose. It is, hence, plausible that the aggregation of amylose caused by the addition of salts would play an important role in the formation of the stable amylose-iodine complex.

Since amylose has been known to have few ionized groups (about one phosphate group per three thousand glucose residues, calculated from the data obtained by Kerr et al.[16]), adsorbed iodide ions would mainly be the source of the charges that prevent the aggregation of amylose.

Although the nature of the amylose-iodine interaction remains controversial[17], Stein and Rundle[3] have presented a theory for the dipolar attraction between amylose and iodine in solid state on the basis of the helical model of amylose. It should be noted that, according to them, the dipolar field provided from the helical amylose in aqueous solution may be

16) R. W. Kerr and G. M. Severson, *J. Am. Chem. Soc.*, 65, 193 (1943).

17) H. Murakami, *J. Chem. Phys.*, 22, 367 (1954); C. T. Greenwood and H. Rossotti, *J. Polymer Sci.*, 27, 481 (1958).

materially reduced by polarization of water, and no strong interaction between amylose and iodine may occur if the helical amylose is surrounded by water molecules with a higher macroscopic dielectric constant.

It is not unreasonable to thnik that since the complexes consisting of amylose, iodine and iodide ion are negatively charged, they can be kept separated without aggregation owing to their Coulombic repulsive force, if the concentration of cation is largely insufficient to repress the force. The complex thus formed may be supposed to be meta-stable since amylose-iodine interaction may not be strong for the reason described above, and may be called intermediate complex. The intermediate complex may not have well-shaped helical structure because of the charges due to adsorbed iodide ion.

By addition of salts these intermediate complexes may tend to gather to aggregates through which water may be excluded from the interstices among them. Moreover, if the aggregates contain helical amylose molecules in close packing, it may be expected that the dipolar field in a given helix would be enhanced by surrounding amylose chains. Then the amylose-iodine complexes formed in aqueous solution would be stabilized by the formation of the aggregates in which the helical complexes would probably be refined on shape and oriented parallel to one another.

Foster and Paschall[18] found that the disaggregation of amylose aggregates results in a decrease in iodine binding capacity and a decrease in the wavelength of maximum absorption of the complex. They also suggested the significance of crystalline order in the complex formation. This seems to support the authors' opinion concerning the nature of the complex as above described. Mayer and Bernfeld[19] considered that the adsorption of iodine occurs on colloidal micells of amylose molecules. Although their conclusion may not be accepted without modification from the helical theory generally accepted at present, their experimental results also indicate that the state of the aggregate of amylose may be closely related to the formation of the blue complex.

In conclusion, it may be said that the blue complex can hardly be formed in solution where ionic strength is so low that the aggregation of the complex takes place with difficulty. The stable state of the blue amylose-iodine complexes formed in solution may be parallel-oriented aggregates composed of helical complexes.

Summary

The relations between the light absorptions of the amylose-iodine complex and the concentrations of various added salts have been investigated. The iodide ion appears to be indispensable for the formation of the blue complex. The complex formation can be greatly aided by certain salts when the concentration of iodide is low. This salt effect suggests that the parallel-oriented aggregation of amylose-iodine helical complex is essential for the formation of the stable blue complex.

The authors wish to thank Miss Setsuko Sugawara for her technical assistance in this study.

Laboratory of Biophysical Chemistry
College of Agriculture
University of Osaka Prefecture
Sakai, Osaka

18) J. F. Foster and E. F. Paschall, *J. Am. Chem. Soc.*, **75**, 1181 (1953).

19) K. H. Meyer and P. Bernfeld, *Helv. Chim. Acta*, **24**, 389 (1941).

Amylose-Iodine Complex. III. Potentiometric and Spectrophotometric Studies

By Takashi KUGE and Sôzaburo ONO

(Received March 21, 1960)

Potentiometric titrations and spectrophotometric measurements have been used as methods of studying the starch-iodine reaction by many authors. Most of these investigations, however, were carried out at room temperature. When the temperature of amylose solution is raised, it may be expected that the amylose molecule enormously increases its flexibility and takes a more random configuration[1]. In parallel with this, it becomes difficult to form amylose-iodine complex with increasing temperature. Hence, it seems interesting to investigate the temperature dependence of the amylose-iodine reaction.

In the present paper, the authors describe the enthalpy change of the amylose-iodine reaction obtained by the potentiometric and/or the spectrophotometric method at wider temperature range than previously reported, and some information concerning the nature of the complex which is deduced from the results obtained.

Experimental

Potentiometric Titrations with Iodine.—The amyloses used for potentiometric titrations were those obtained from potato (white potato), corn and sweet potato starch by hot water extraction[2] at 60°, 69° and 72°C, respectively.

The technique of potentiometric titration was essentially similar to that described by Bates et al.[3]: In our procedure about 0.015 g. of amylose was dispersed in 10 ml. of 0.5 M potassium hydroxide. When dispersal was complete, the mixture was made neutral to methyl orange by the addition of hydrochloric acid, added 10 ml. of 0.5 M potassium iodide, and then filled up to 100 ml. The solution thus prepared was titrated with 0.0005 M iodine in 0.05 M potassium iodide solution. A control solution which contains the same components as that described above but no amylose, was also titrated in the same way.

Titration was carried out in a cell immersed in a thermostat.

For the potential measurements, a Yokogawa high precision potentiometer was used in conjunction with a normal calomel electrode and a bright platinum electrode.

Spectrophotometric Measurements.—The amylose used was obtained from potato starch by a selective precipitation method with *n*-butyl alcohol and isoamyl alcohol[4].

Spectrophotometric titration was carried out as follows: 100 mg. of amylose was dispersed in 100 ml. of 0.5 M potassium hydroxide and diluted to 500 ml. with distilled water after neutralization with hydrochloric acid. Each 25 ml. of the solution was diluted to 100 ml., a requisite amount of 0.1 M potassium iodide being added to make the final solution 0.01 M in iodide after the addition of 1 to 22 ml. of a solution of 0.0005 M iodine in 0.001 M potassium iodide. Although the absorptions of the sample in which 6.5 ml. or less of 0.0005 M iodine solution was contained, were directly measured, those of the sample in which 8.0 ml. or more of the iodine solution was contained, were measured after twofold dilution with water. The absorptions were measured at 15°C at wavelengths of 288, 350 and 640 mμ, respectively.

The change of the absorption with temperature change was measured in the following way: The solution (0.005% amylose, 2.5×10^{-5} M iodine), in which the amount of iodine was nearly half of the amount to saturated amylose, was prepared in the same way as that for the spectrophotometric titrations. This solution was placed in a 250 ml. colored volumetric flask and kept in a thermostat at each given temperature for 20 min. and then aliquots of the solution were used to measure the absorption at wavelengths of 288, 350 and 640 mμ, respectively.

A Shimadzu type QB-50 spectrophotometer with thermo-regulating device was used with 1 cm. thick silica cell throughout this investigation.

Results and Discussion

Potentiometric Titrations.—Titration curves of potato amylose and control solution (0.05 M in potassium iodide) at various temperatures are shown in Fig. 1. It can be seen that the potential of these systems becomes high with increasing temperature. At a given temperature the potential is determined by the concentration of free iodine which is in equilibrium with amylose-iodine complex, since the concentration of iodide ion is kept practically constant.

Referring to the results of Dube[5] and Gilbert et al.[6], it may be considered that the ratio of

1) S. Ono, T. Kuge and Y. Yoshikawa, presented at the 11th Symposium on Colloid Chemistry held in Sendai, September, 1958.

2) K. G. Krishnaswamy and A. Sreenivasan, *J. Biol. Chem.*, **176**, 1253 (1948).

3) F. L. Bates, D. French and R. E. Rundle, *J. Am. Chem. Soc.*, **65**, 142 (1943).

4) E. J. Wilson, T. J. Schoch and C. S. Hadson, ibid., **65**, 1380 (1943).

5) H. A. Dube, Dissertation, Iowa State College, (1947), through "Chemistry and Industry of Starch", edited by R. W. Kerr, Academic Press, Inc., New York, (1950), Chapter XVII, p. 461.

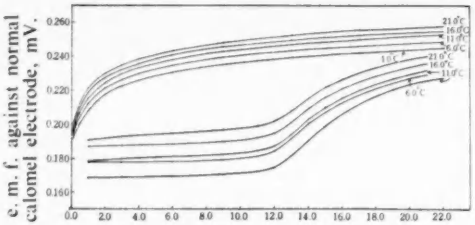


Fig. 1. Potentiometric titration curves: 100 ml. of 0.012% potato amylose solution (the lower five curves), 100 ml. of 0.05 M potassium iodide (the upper five curves).

adsorbed iodide ions to adsorbed iodine molecules on amylose is unity in the present experimental conditions (0.05 M in potassium iodide). Therefore, the amylose-iodine reaction can be expressed as

$$\mathrm{Am} + n\mathrm{I}_2 + n\mathrm{I}^- \rightleftarrows \mathrm{Com} \qquad (1)$$

where Am and Com represent amylose and amylose-iodine complex, respectively.

The equilibrium constant K, which is constant for the amylose of a given chain length, is then

$$K = [\mathrm{Com}] / [\mathrm{Am}]\,[\mathrm{I}_2]^n\,[\mathrm{I}^-]^n$$

hence

$$K^{1/n} = [\mathrm{Com}]^{1/n} / [\mathrm{Am}]^{1/n}\,[\mathrm{I}_2]\,[\mathrm{I}^-] \qquad (2)$$

In the amylose-iodine reaction, the activities of the complex and amylose can be regarded as unity, in the manner of a heterogeneous reaction. Eq. 2, therefore, may be written as

$$K^{1/n} = 1 / [\mathrm{I}_2]\,[\mathrm{I}^-] \qquad (3)$$

Since the potential of the horizontal part of the titration curve is lower than the lowest potential obtained with the control solution in a given temperature, as can be seen in Fig. 1, direct calculation of free iodine from the potential, which was made by Bates et al.[3], is impossible in the present case. Accordingly, the calculation of $K^{1/n}$ was carried out as follows. The potential difference, E_d, between the potential of the sample at the mid-point of the titration (i.e. characteristic potential[7]) and that of the control solution at the corresponding titration point, can be written as

$$E_d = (RT/2F) \ln ([\mathrm{I}_2]_c / [\mathrm{I}_2]_s) \qquad (4)$$

where $[\mathrm{I}_2]_c$ is the concentration of free iodine in the control solution and $[\mathrm{I}_2]_s$ corresponds to that of the sample, and R and T are gas constant and absolute temperature, respectively, and F is Faraday constant. From Eqs. 3 and 4

$$E_d = (RT/2F) \ln [\mathrm{I}_2]_c\,[\mathrm{I}^-]_s\,K^{1/n} \qquad (5)$$

or

$$E_d = (RT/2F) \ln ([\mathrm{I}_3^-]_c\,K^{1/n}/K_1) \qquad (6)$$

where $[\mathrm{I}^-]_s$ is the concentration of iodide ion in the sample, $[\mathrm{I}_3^-]_c$ is the concentration of triiodide ion in the control solution and K_1 is the equiliblium constant of the following reaction

$$\mathrm{I}_2 + \mathrm{I}^- = \mathrm{I}_3^-$$

Therefore, $K^{1/n}$ can be calculated from E_d, $[\mathrm{I}_3^-]_c$ and K_1 which is given in literature[8], using Eq. 6.

In Table I are listed observed $K^{1/n}$, the enthalpy change of the reaction obtained from the temperature dependence of $K^{1/n}$, characteristic potentials, intrinsic viscosities of amylose in 1 M potassium hydroxide by Ostwald viscometer[9] and the temperature of water extraction of amylose. As the temperature of the water extraction is raised, the mean length of the extracted amylose chain may be expected to become longer. This is inferred from their characteristic potentials[7] and intrinsic viscosities, though the contamination of amylopectin may increase. The enthalpy changes of the reaction with amyloses from different sources agree within experimental error in spite of the difference of the size of molecule among these amyloses. Similar values have also been obtained with amyloses partially degraded by bacterial α-amylase.

Since the results obtained at temperatures higher than those described above are confused owing to the vaporization of iodine during the procedure, the analysis of the results is not made as those obtained at a lower temperature. However, it seems worth while to mention that the end-point of titration shifted to higher concentration of iodine and became obscure with increasing temperature, from which the occurrence of a complex having a short polyiodine-chain (also suggested in the following spectrophotometric studies) at a higher temperature might be supposed.

Spectrophotometric Studies.—Fig. 2 shows the results obtained from spectrophotometric titrations at 15°C. At the beginning of the titration, the curve for 640 mμ has a small lag and thereafter increases linearly with the amount of iodine added until the vicinity of end-point.

6) G. A. Gilbert and J. V. R. Marriott, *Trans. Faraday Soc.*, 44, 84 (1948).

7) J. F. Foster, Dissertation, Iowa State College, Ames, Iowa (1943), through "Chemistry and Industry of Starch", edited by R. W. Kerr, Academic Press, Inc., New York, (1950), Chapter XVII, p. 460.

8) M. Davies and E. Gwynne, *J. Am. Chem. Soc.*, 74, 2748 (1952).

9) Viscometric studies were carried out in a way similar to that described by S. Lansky et al., *J. Am. Chem. Soc.*, 71, 4066 (1949).

TABLE I. SOME CONSTANTS OBTAINED WITH AMYLOSES FROM DIFFERENT SOURCES

Amylose	Temp. of water extraction	Equilibrium constant $K^{1/n}$ at 16°C	Enthalpy change kcal./mol. iodine	Charact. potential v.s. N. C. E. at 16°C	Intrinsic viscosity
Potato	60°C	2.2×10^9	−15.5	0.1890	0.80
Corn	69°C	2.4×10^9	−15.8	0.1878	1.20
Sweet potato	72°C	3.0×10^9	−15.3	0.1856	2.15

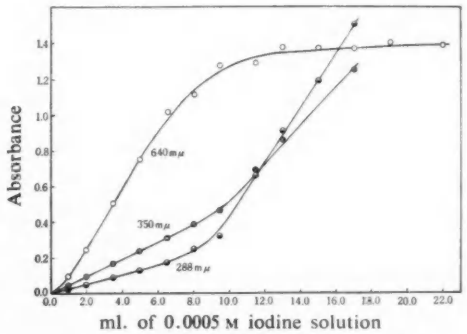


Fig. 2. Spectrophotometric titration curves: 100 ml. of 0.005% potato amylose solution.

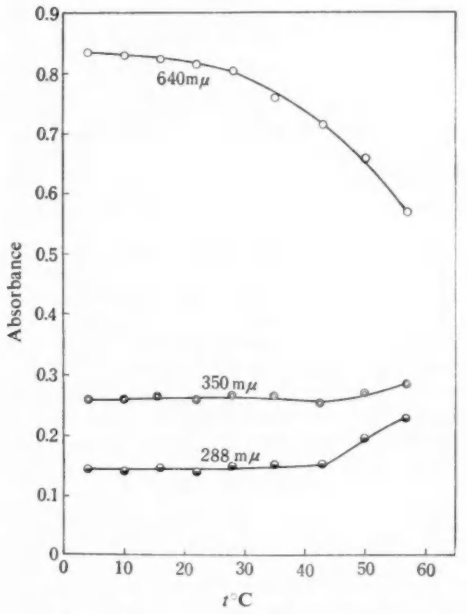


Fig. 3. Temperature dependences of absorptions at wavelengths of 640, 350 and 288 mμ.

The absorptions of ultraviolet region (at 288, 350 mμ) appear to have no such lag and their sharp increases due to free triiodide ion are seen after the end-point. Using the slopes of the initial linear parts of the titration curves, the molecular extinction coefficients of iodine in the complex were estimated to be 36300, 10500 and 5230 at wavelengths of 640, 350 and 288 mμ, respectively. The latter two values are in good agreement with that calculated from Mould's report[10] by the present authors.

The temperature dependences of absorptions are shown in Fig. 3, in which it can be seen that the absorption at 640 mμ decreases in a regular fashion with increasing temperature, while the absorption of the ultraviolet region (at 288 and 350 mμ) remains almost unchanged and has rather a tendency to increase at temperature higher than 40°C.

Free triiodide ion in an aqueous solution exhibits two ultraviolet absorption bands whose maxima are at 288 and 353 mμ[11], respectively. Since the concentration of free iodine molecules which are in equilibrium with the complexes may increase with increasing temperature and consequently the concentration of free triiodide ions may increase, the above described temperature dependence of absorption in the ultraviolet region may be principally ascribed to the increase of the free triiodide ions.

The absorption at 640 mμ may be considered to be due to the complex alone, since free iodine concentration is low enough that its contribution to the absorption can be neglected. Assuming that the absorption of the ultraviolet region is determined by the concentrations of the complex and free triiodide ion, as a first approximation, the following equations can be derived.

$$E_{640} = C_c\varepsilon_{c640} \tag{7}$$

$$E_{288} = C_c\varepsilon_{c288} + C_i\varepsilon_{i288} \tag{8}$$

$$E_{350} = C_c\varepsilon_{c350} + C_i\varepsilon_{i350} \tag{8'}$$

where E represents the absorbance observed, C_c and C_i are the concentrations of the complex and free triiodide ion, respectively, ε_c and ε_i are the molecular extinction coefficients of the complex and free triiodide ion, respectively, and the numerical suffixes represent their corresponding wavelengths in mμ. The concentrations of the complex, for the sake of convenience, are expressed as moles of iodine bound.

10) D. L. Mould, *Biochem. J.*, **58**, 593 (1954).
11) A. D. Awtrey and R. E. Connick, *J. Am. Chem. Soc.*, 73, 1842 (1951).

TABLE II. TEMPERATURE DEPENDENCE OF SOME PROPERTIES OF POTATO AMYLOSE

Temp. t C	Concn. of complex C_c mol./l.	Concentration of free triiodide ion C_{1288} mol./l.	C_{1350} mol./l.	C_{1c}	Concn. of short polyiodine C_u mol./l.	Equilibrium constant $K_s^{1/n}$	$K^{1/(n-s)}$	Enthalpy change kcal./mol. iodine
1.0	—	—	—	—	—	—	9.2×10^{9}	
4.0	2.29×10^{-5}	5.8×10^{-7}	6.2×10^{-7}	5.6×10^{-7}	1.5×10^{-7}	2.4×10^{9}	—	
6.0	—	—	—	—	—	—	4.8	
10.0	2.28	5.6	6.9	4.9	5.2	2.3	—	
11.0	—	—	—	—	—	—	4.2	
16.0	2.27	6.1	9.2	4.6	1.1×10^{-6}	2.1	2.2	−15.5[a]
21.0	—	—	—	—	—	—	1.3	
22.0	2.24	5.3	9.2	3.4	1.4	2.4	—	
28.0	2.22	8.1	1.31×10^{-6}	5.6	1.8	1.3	—	
35.0	2.09	1.06×10^{-6}	1.81	6.9	2.8	9.1×10^{-1}	—	
43.0	1.97	1.24	1.88	9.2	2.4	6.0	—	
50.0	1.81	2.51	3.12	2.21×10^{-6}	2.3	2.2	—	−20.8
57.0	1.57	3.65	4.62	3.17	3.6	1.4	—	

a) This value was obtained by potentiometric titration.

Assuming that these molecular extinction coefficients are independent of temperature, the concentration of the free triiodide ion at each temperature may be found, using Eqs. 7 and 8 or 8′ [12]. The values obtained are shown in Table II, in which C_{1288} and C_{1350} are the concentrations of free triiodide ion estimated by using the absorbance observed at wavelengths of 288 and 350 mμ, respectively. It can be seen that C_{1288} and C_{1350} do not agree with each other and that the former is smaller than the latter at any temperatures. This seems to suggest that Eq. 8 or 8′ would not be an adequate expression for the observed absorbance. A more favorable equation than Eq. 8 or 8′ is seemed to be written as

$$E=C_c\varepsilon_c+C_i\varepsilon_i+C_u\varepsilon_u+C_a\varepsilon_a \qquad (9)$$

where C_u represents the concentration of the polyiodine of short length and is also expressed as moles of single iodine molecule, ε_u is the molecular extinction coefficient of iodine in the polyiodine, C_a and ε_a represent the concentration and molecular extinction coefficient of the triiodide ion loosely adsorbed on the surface of amylose, respectively. It would appear that little attention has been given to the ultraviolet absorption of the amylose-iodine complex and an investigation of the ultraviolet absorption has been inconclusive[10,13]. However, the fact that the titration curve at 640 mμ has a small lag may reasonably suggest the existence of polyiodine of short length which exhibits no appreciable absorption at 640 mμ but ultraviolet absorption similar to that of the complex. It may be assumed that the absorption spectrum of the triiodide ion is unchanged whether in the adsorbed state described above or in an aqueous solution.

The use of Eq. 8 or 8′ instead of Eq. 9 would cause an error ΔC_i being always positive, in the estimation of the concentration of the free triiodide ion. The error can be expressed as follows.

$$\Delta C_i=C_u\varepsilon_u/\varepsilon_i+C_a\varepsilon_a/\varepsilon_i \qquad (10)$$

Since, ε_u may be roughly assumed to be not appreciably different from ε_c and ε_a may be considered as equal to ε_i, the ratio of ε_u to ε_i depends on wavelength while that of ε_a to ε_i seems to be unity, independent of wavelength. The smaller the ratio of ε_u to ε_i, the smaller the error becomes. Therefore, C_{1288} should be smaller than C_{1350}, and the experimental results given in Table II satisfy this requirement.

The true concentration of free triiodide ion, C_{it}, would be written as

$$C_{it}=C_{1288}-C_u\varepsilon_{u288}/\varepsilon_{1288}-C_a\varepsilon_{a288}/\varepsilon_{1288} \qquad (11)$$

or

$$C_{it}=C_{1350}-C_u\varepsilon_{u350}/\varepsilon_{1350}-C_a\varepsilon_{a350}/\varepsilon_{1350} \qquad (11')$$

Therefore, C_u can be obtained by subtraction of Eq. 11 from Eq. 11′. C_u obtained thus is shown in Table II, and the increase of C_u with increasing temperature is seen. C_{1c} listed in Table II is the corrected concentration of free triiodide ion in respect to C_u. It should be noted that C_{1c} is not free from the error due to C_a. C_{1c} was used to calculate $K^{1/n}$, and the

12) ε_{1288} and ε_{1350} used in the calculation were 40000 and 26000, respectively, which were obtained from our experiments and have no essential difference from the values in reference 11.

13) Y. Tanizaki, T. Kobayashi and N. Andô, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **80**, 445 (1959).

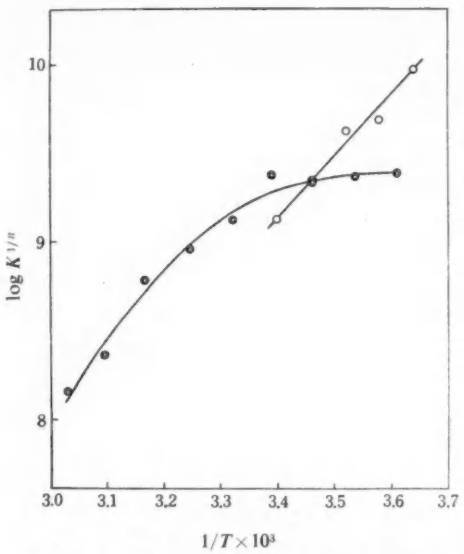


Fig. 4. $\log K^{1/n}\sim 1/T$ plot for potato amylose: —◎—◎—, $\log K_s^{1/n}$, obtained from spectrophotometric method; —○—○—, $\log K^{1/n}$, obtained from potentiometric titration.

value obtained is given in Table II expressed as $K_s^{1/n}$. In Fig. 4, $\log K_s^{1/n}$ is plotted against reciprocal absolute temperature, and it can be seen that the slope of the curve becomes steeper with increasing temperature. Comparing $K_s^{1/n}$ with $K^{1/n}$ obtained by the potentiometric titrations, it can be found that they have similar values at about 16°C, but the slopes as above described, from which the enthalpy change of the reaction would be calculated, are quite different from each other in the neighborhood of the temperature. Such difference may be reasonably explained in terms of the error in the spectrophotometric estimation of free triiodide ion concentration at low temperature region. Considering the data obtained by the potentiometric titrations, it can be seen that free iodine concentration in equilibrium with the complex should decrease with decreasing temperature. However, C_{Ie} rather increases with decreasing temperature in the temperature region lower than 22°C. This fact suggests that the amount of the triiodide ion adsorbed on the surface of amylose may be comparable with that of free triiodide ion in the temperature region below 22°C, and a serious error is caused in the estimation of free triiodide ion concentration. It may also be supposed that the lower the temperature, the more the contribution of C_n to C_{Ie} becomes. Therefore, it can be inferred that $\log K_s^{1/n}\sim 1/T$ plot should be steeper than that seen in Fig. 4, especially in the lower temperature region. Consequently, the equilibrium constant obtained spectrophotometrically should be larger than that obtained potentiometrically. This result may not be unreasonable, since $K^{1/n}$ is considered as a function of chain length of amylose, being larger with increasing chain length of amylose[9,14], and amylose used in the spectrophotometric study is thought to be longer than that used in potentiometric titration. The selective precipitation method of amylose seems to give amylose with longer chain length than that obtained by the hot water extraction.

For the reason described above, the spectrophotometric method may not be applicable for the estimation of free iodine concentration at room temperature, at which potentiometric titrations are adequate. However, with increasing temperature, when the concentration of free iodine in equilibrium with the amylose-iodine complex increases, the concentration of the free iodine appears to be reasonably estimated by the spectrophotometric method, since, at the higher temperature, C_n can be neglected without a serious error, compared with the free iodine concentration. Even when the vaporization of iodine takes place, free iodine concentration may remain approximately constant as far as amylose-iodine complex exists in solution. Therefore, the present results obtained at a relatively high temperature region may be available for the discussion of the temperature dependence of the amylose-iodine reaction.

From the inspection of the slope of $\log K_s^{1/n}\sim 1/T$ plot in Fig. 4, the enthalpy change of the reaction appears to increase with increasing temperature. The enthalpy change obtained at 50°C is −20.8 kcal. Comparing this value with that obtained potentiometrically at 16°C, −15.5 kcal., it seems certain that the enthalpy change increases as the temperature is raised. This may also be checked with the fact that all the enthalpy changes of the reaction hitherto obtained at different temperatures below 35°C by other authors[5,6,15,16] are −11 to −19 kcal.

At the temperature at which the color of the complex disappears, the free energy change of the reaction should be nearly zero so that the entropy change should have a large value almost corresponding to the enthalpy change at the temperature. This seems to indicate that the amylose, which is free from iodine, would come to take more random configurations

14) Also seen in Table I.

15) Presented at the 12th Symposium on Structural Chemistry held in Fukuoka, November, 1957, by H. Azumi et. al.

16) J. Holló and J. Szejtli, *Periodica Polytechnica*, **1**, 223 (1957).

with increasing temperature through the rupture of both inter- and intramolecular hydrogen bonding. This thought is also supported by the results obtained from the studies in viscosity and optical rotatory power[1].

Summary

The amylose-iodine reaction has been studied in a wider temperature range than hitherto reported by both potentiometric and spectrophotometric methods.

The enthalpy change of the reaction obtained potentiometrically is −15.5 kcal. at 16°C, and appears to be independent of the chain length and the kind of parent starch of amylose.

The change of the absorptions of the amylose-iodine complex with temperature has been pursued at wavelengths of 650, 288 and 350 mμ, respectively.

The triiodide ion adsorbed on the surface of amylose causes an error in the spectrophotometric estimation of free iodine concentration. At the higher temperature, however, this error may be neglected, and the results obtained seem to be reliable.

The enthalpy change appears to increase with increasing temperature and −20.8 kcal. has been obtained at 50°C.

Laboratory of Biophysical Chemistry
College of Agriculture, University of
Osaka Prefecture, Sakai, Osaka

Spectrochemical Study of Microscopic Crystals. XXII.[1] *The Structure of Cupric Mono-, Di- and Trichloroacetate*[2]

By Shoichiro YAMADA, Hiroaki NISHIKAWA and Ryutaro TSUCHIDA

(Received April 27, 1960)

Previously the present authors[3,4] studied the structure of cupric *n*- and *iso*-alkanoates, anhydrous and hydrated, and showed mostly on the basis of the polarized absorption spectra that these carboxylates, in the crystalline state and in non-coordinating organic solvents, consist of binuclear molecules which are similar to that of cupric acetate and have a sort of copper-to-copper linkage within the molecules. In continuation of the former study, the structure of cupric mono-, di- and trichloroacetate has now been examined, and the effect of replacement by chlorine atoms discussed in relation to the dimer formation of the cupric carboxylates. A brief account of a part of this study was formerly presented as a short communication[5], the details of which are given below.

Experimental

Materials. — All the cupric chloroacetates were prepared according to the method of Bateman et al.[6,7] from cupric carbonate and the corresponding chloroacetic acid or from sodium carboxylate and cupric sulfate. All the compounds were identified by elemental analysis. Crystals of the anhydrous forms of the compounds were obtained by recrystallization using non-aqueous solvents after dehydrating the hydrated crystals. The absence of the water of crystallization was confirmed by infrared spectra and elemental analysis.

Anhydrous and hydrated cupric monochloroacetate were obtained in green crystals. The hydrated monochloroacetate corresponds to the formula, $Cu(ClCH_2COO)_2 \cdot 2.5H_2O$. Calcd.: Cu, 21.5; Cl, 30.6; C, 20.7; H, 2.61. Found: Cu, 21.5; Cl, 30.4; C, 20.4; H, 2.55%. A blue form has also been prepared besides the green form. Thus the blue crystals of the anhydrous monocloroacetate were obtained, when ethanol was added to an aqueous solution of the monochloroacetate.

Anhydrous cupric dichloroacetate was prepared in green crystals. Hydrated crystals, which are colored blue, were also obtained together with the anhydrous crystals, when recrystatallization was made from water. The result with the hydrated crystal is reported elsewhere.

Cupric trichloroacetate was obtained in blue crystals after recrystallization from water. The

1) Part XXI of this series, S. Yamada, H. Nishikawa and R. Tsuchida, This Bulletin, **33**, 930 (1960).
2) Presented in part at the Symposium on Co-ordination Copounds, Nagoya, November, 1958.
3) R. Tsuchida and S. Yamada, *Nature*, **176**, 1171 (1955). R. Tsuchida, S. Yamada and H. Nakamura, ibid., **178**, 1192 (1956).
4) S. Yamada, H. Nakamura and R. Tsuchida, This Bulletin, **30**, 953 (1957); ibid., **31**, 303 (1958).
5) R. Tsuchida, S. Yamada and H. Nakamura, *Nature*, **181**, 479 (1958).
6) W. G. Bateman and A. B. Hoel, *J. Am. Chem. Soc.*, **36**, 2517 (1914).
7) W. G. Bateman and D. B. Conrad, ibid., **37**, 2553 (1915).

elemental analysis by the present authors indicates that the blue crystals correspond to the formula, $Cu(Cl_3C\cdot COO)_2\cdot 3H_2O$, in agreement with the analysis by Bateman et al.[6]

Measurements.—Quantitative dichroism measurements in the visible and ultraviolet regions were carried out at room temperature by Tsuchida and Kobayashi's microscopic method[8].

Absorption spectra of the compounds in solution were determined by a Beckman DU spectrophotometer. It was confirmed in all the measurements that Beer's law was obeyed in the presence of the free carboxylic acid.

The notations in the present paper are the same as close used in the former papers of this series. Extinction coefficients, ε, are calculated on the basis of a formula-weight instead of a molecular weight.

Results and Discussion

Structure of Cupric Chloroacetates in the Crystalline State. — Previously the present authors[3-5] reported about the structure of cupric formate, acetate and higher carboxylates on the basis of their polarized absorption spectra. A similar treatment will be made in the present article concerning the structure of cupric chloroacetates. Since some of the results and the discussions in the former papers are used in the present work, it would be appropriate to give a brief account of the conclusions obtained formerly.

Cupric salts of acetic acid and higher fatty acids in the crystalline state and in non-coordinating organic solvents show in the visible and the near-ultraviolet region a band at about $80\times10^{13}\,sec^{-1}$ (band II) besides a band at about $43\times10^{13}\,sec^{-1}$ (band I), whereas most cupric compounds of an ordinary type show in this wavelength region only one band which may be considered as corresponding to the band at about $43\times10^{13}\,sec^{-1}$ of cupric alkanoates. With the cupric carboxylates consisting of binuclear molecules, the following features characteristic of these compounds were observed in the dichroism. "For the band I, the electric vector is more strongly absorbed along the complex plane than along the *Z*-direction of the binuclear molecule, the *Z*-direction being the direction connecting the two copper ions in the binuclear molecule. For the band II, the electric vector is more strongly absorbed along the *Z*-direction. The polarization for the band II is thus seen to be the reverse of the polarization for the band I". These facts are considered to show that the compounds involve along the *Z*-direction a sort of copper-to-copper linkage characteristic of the binuclear molecules of the cupric alkanoates. On the contrary, cupric formate tetrahydrate[3,4], which is known from examination with X-rays[9] to have no copper-to-copper linkage[10], shows no absorption band at about $80\times10^{13}\,sec^{-1}$. Thus the appearance of the band at about $80\times10^{13}\,sec^{-1}$, combined with the polarization property of the band, may be regarded as showing the existence of the binuclear molecule similar to that of cupric acetate hydrate involving the copper-to-copper linkage.

The results of the present dichroism measurements are shown in Figs. 1 to 4. The components of the cupric compounds in the crystalline state may be under rather weak perturbation owing to the crystalline field, but the polarized absorption spectra as determined on a suitable plane are regarded as representing the main features of the dichroism of the component molecule. The exact arrangement of the molecules in the crystals is unknown for all the cupric compounds reported in the present work. It is easy, however, to find the principal characteristics of the dichroism of the molecules themselves by comparing their polarized absorption spectra with those of cupric acetate monohydrate, whose crystal structure was accurately determined, and also with those of cupric alkanoates, which were extensively studied on the basis of the comparison with cupric acetate.

The present measurement, as seen in Fig. 1, indicates that the green anhydrous cupric monochloroacetate in the crystalline state shows an absorption band at about $80\times10^{13}\,sec^{-1}$, in

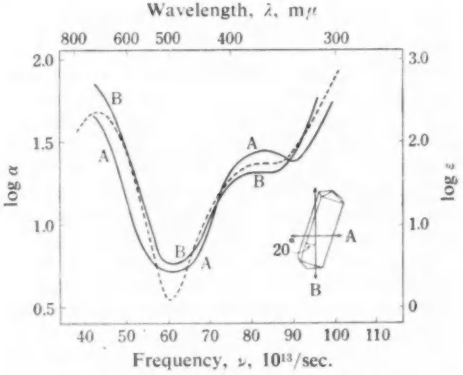


Fig. 1. Absorption spectra of anhydrous $Cu(Cl\cdot CH_2COO)_2$ in the crystalline state (——) and in dioxane (----).

8) R. Tsuchida and M. Kobayashi, "*The Color and the Structure of Metallic Compounds*", (*Kinzokukagōbutsu no Iro to Kōzō*), Zoshindo, Osaka, 1944, p. 180 (in Japanese).

9) R. Kiriyama, H. Ibamoto and K. Matsuo, *Acta Cryst.*, 7, 482 (1954).

10) Cupric formate in dioxan exists mostly as binuclear molecules of the acetate-type, and crystals of the formate consisting of the binuclear molecules were isolated. Cf. R. L. Martin and H. Waterman, *J. Chem. Soc.*, 1959, 1359.

addition to a band at about $40\times10^{13}\,sec^{-1}$. Moreover, it is recognizable that the polarization for the former band is found to be the reverse of the polarization for the latter. This relationship, which closely resembles that encountered in the case of the acetate monohydrate and higher alkanoates[3,4], may suggest that the green anhydrous cupric monochloroacetate in the crystaline state consists of the binuclear molecules of the acetate-type, since the dichroism of the anhydrous monochloroacetate may be readily understandable when it is assumed that A- and B-absorption in Fig. 1 represent the main characteristics of the *Z*- and *X*, *Y*-absorption of the binuclear molecule, respectively.

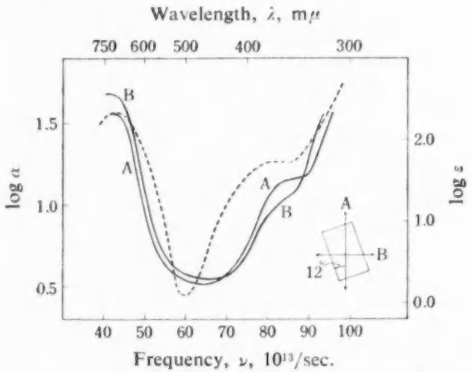


Fig. 2. Absorption spectra of $Cu(Cl\cdot CH_2COO)_2\cdot 2.5H_2O$ in the crystalline state (——) and in dioxane (----).

The present measurement also indicates that cupric monochloroacetate hydrate in the crystalline state shows dichroism with characteristics similar to those of cupric acetate, as is readily seen in Fig. 2. Thus the monochloroacetate hydrate shows an absorption band at about $80\times10^{13}\,sec^{-1}$ in addition to a band at about $40\times10^{13}\,sec^{-1}$. The main characteristics of the dichroism with the component molecule may be recognized, although the dichroism is not very remarkable. It is, therefore, concluded that cupric monochloroacetate hydrate consists of the binuclear molecule of the acetate-type. As comparison of the dichroism of the green anhydrous monochloroacetate with the dichroism of the green monochloroacetate hydrate shows, the presence of the water molecules of crystallization is expected to have a more or less remarkable effect upon the copper-to-copper linkage, but no marked difference was observed in the absorption maxima in the visible and the near-ultraviolet regions between the anhydrous and the hydrated form.

The blue form of anhydrous cupric monochloroacetate is also obtained when ethanol is added to an aqueous solution of the monochloroacetate. The blue crystal is found in the present work to show an absorption band in the near-infrared region, but no absorption band at about $80\times10^{13}\,sec^{-1}$. It is, therefore, concluded that the blue form of the monochloroacetate contains no binuclear molecule.

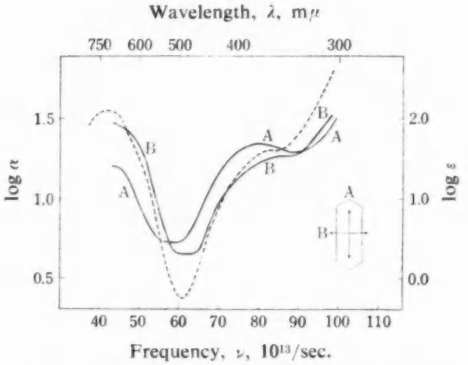


Fig. 3. Absorption spectra of anhydrous $Cu(Cl_2\cdot CHCOO)_2$ in the crystalline state (——) and in dioxane (----).

The present measurement, as seen in Fig. 3, indicates that the green crystal of anhydrous cupric dichloroacetate shows a band at about $80\times10^{13}\,sec^{-1}$ besides a band at about $40\times10^{-13}\,sec^{-1}$, the polarization of the former band being the reverse of the polarization of the latter. This relationship, which is similar to the relationship with cupric acetate hydrate, indicates that anhydorus cupric dichloroacetate in the crystalline state consists of binuclear molecules.

Cupric dichloroacetate is known to exist in a tetrahydrate as well. The dichroism of this form, which is colored blue, is not identical with that of the anhydrous form. The structure of the tetrahydrate will be discussed separately in a later communication.

Cupric trichloroacetate is obtained in the tetrahydrate, when recrystallized from water, ethanol-water, or dioxane-water. The trichloroacetate tetrahydrate shows the dichroism which is quite different from the dichroism of the green form of the mono- or the dichloroacetate. Thus the trichloroacetate tetrahydrate shows only one absorption band at about $40\times10^{13}\,sec^{-1}$, but no absorption band in the neighborhood of the wavelength region around $80\times10^{13}\,sec^{-1}$, as seen in Fig. 4. The absorption spectrum of the trichloroacetate, which is similar to the absorption spectra of ordinary cupric complexes such as cupric sulfate, indicates clearly that cupric trichloroacetate

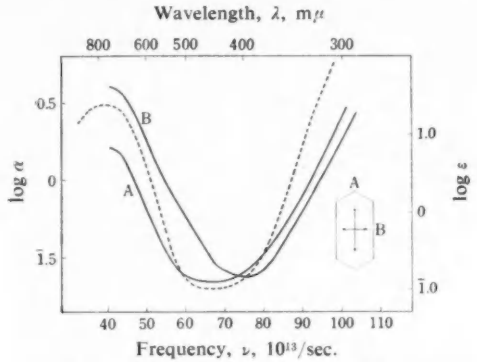


Fig. 4. Absorption spectra of $Cu(Cl_3C\cdot COO)_2\cdot 4H_2O$ in the crystalline state (——) and in dioxane (----).

tetrahydrate contains no binuclear molecule of the acetate-type in the crystalline state.

Anhydrous cupric trichloroacetate, which is obtained in green crystals when recrystallized from ethyl ether after dehydration of the tetrahydrate, may probably consist of the binuclear molecule of the acetate-type, but the dichroism measurement with the crystal has not been performed in the present work. As discussed in a later paragraph of the present article, the ultraviolet spectrum shows that the anhydrous trichloroacetate exists as the binuclear molecules in ethyl ether.

Magnetic susceptibility at room temperature was reported with cupric mono-, di- and trichloroacetates in the solid state[11]. The results seem to be compatible with the conclusions from the present study about the structure of these compounds.

Structure of Cupric Chloroacetates in Solution.—The binuclear cupric carboxylates in water are found to undergo dissociation most readily. Even cupric acetate exists in this solvent mostly as the monomer in equilibrium with the binuclear form in a smaller proportion, although the degree of dissociation depends upon the concentration of the solution. All the cupric monocarboxylates so far examined have been found to exist in water as their monomers to a considerable extent if they are soluble in water at all. This may evidently be due to the great tendency of water molecules to be co-ordinated with the cupric ion. On the contrary, dissociation into the monomeric complex-ion has been found to be quite small in non-co-ordinating organic solvents, such as ethanol, dioxane, ethyl ether and chloroform. Needless to say, cupric carboxylates in co-ordinating solvents such as pyridine undergo decomposition to form cupric complexes with the solvent molecules. In the present work, comparison of the absorption spectra is made with the cupric carboxylates in these solvents in almost the same concentration.

It was found in the former investigations that the cupric monocarboxylates having the binuclear molecules of the acetate-type show in non-co-ordinating organic solvents the characteristic absorption spectra with the following features[4].

(1) A band at about $40\times10^{13}\,sec^{-1}$ of the binuclear cupric carboxylate is much stronger than the corresponding band of the mononuclear cupric compounds. (2) Besides the absorption at about $40\times10^{13}\,sec^{-1}$ of the binuclear, cupric carboxylate shows another band at about $80\times10^{13}\,sec^{-1}$, which is about four times less in intensity than the band in the longer wave-length region of the same compound. As reported in the former papers[3,4], these features may be used as criteria in telling whether a cupric carboxylate exists mostly as binuclear molecules of the acetate-type or not.

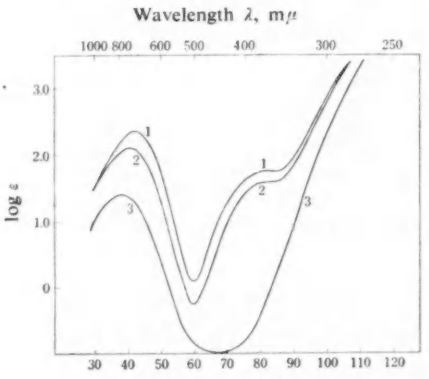


Fig. 5. Absorption spectra of $Cu(Cl_nH_{3-n}C\cdot COO)_2$ in dioxane: 1, $n=1$; 2, $n=2$; 3, $n=3$.

Absorption spectra of cupric chloroacetates in dioxane are shown in Fig. 5. The absorption spectra characteristic of the cupric carboxylates having binuclear structure of the acetate-type are observed with the mono- and the dichloroacetate. Thus the two chloroacetates in dioxan show an absorption band at about $80\times10^{13}\,sec^{-1}$, in addition to a band at about $40\times10^{13}\,sec^{-1}$ which has an extinction coefficient about nine times higher than the corresponding band of an ordinary type of cupric complex such as an aquated ion of copper. This fact indicates, on the basis of the criteria formerly obtained, that the mono- and the

11) M. Kondo and M. Kubo, *J. Phys. Chem.*, **62**, 1558 (1958).

dichloroacetate in dioxane exist, at least in most part, as binuclear molecules.

On the contrary, cupric trichloroacetate in dioxane shows an absorption band at about $40\times10^{13}\,sec^{-1}$, but no absorption band at about $80\times10^{13}\,sec^{-1}$. The spectrum shows the main features of the spectrum of ordinary cupric complexes without specific metal-to-metal linkage. It may be concluded, therefore, that cupric trichloroacetate does not contain the binuclear molecule of the acetate-type in dioxan. It is noteworthy that even anhydrous cupric trichloroacetate in dioxane shows no band at about $80\times10^{13}\,sec^{-1}$, having no binuclear molecule in this solvent. It is most probable that two trichloroacetate ions are coordinated as bidentate ligands with a single cupric ion to form a planar complex-ion.

The present work shows that the maximum for the band I of the cupric carboxylate is displaced bathochromically, when the hydrogen atom is replaced by the chlorine atom in the series from the acetate to the trichloroacetate, as seen in Table I. The fact that this order

Table I. Visible and ultraviolet absorption spectra of cupric monocarboxylates with an empirical formula of $Cu(R\cdot COO)_2$

R	Solvent	Band I		Band II	
		ν	log ε	ν	log ε
CH_3	ethanol	42.8	2.25	82.0	1.65
	dioxane	45.2	2.25	81.1	1.81
$Cl\cdot CH_2$	ethanol	40.2	1.82	82	1.1
	dioxane	42.9	2.34	82	1.74
$Cl_2\cdot CH$	ethanol	38.5	1.58	83	0.5
	dioxane	41.1	2.12	83	1.60
$Cl_3\cdot C$	ethanol	38.0	1.47	none	
	dioxane	38.4	1.44	none	
	ethyl ether	38.3	2.18	ca. 84	1.50

ν: $10^{13}\,sec^{-1}$.

seems to be parallel with the stability of the binuclear molecule may be readily understandable in terms of the ligand field theory, since the band I corresponds to the electronic transitions between the terms formed by splitting from the cupric ion under the ligand field, which is expected to be stronger, as the cupric carboxylates become more stable.

From the above-mentioned observation with the cupric chloroacetates both in the crystalline state and in dioxane, it may be concluded that the tendency of the cupric carboxylates to form binuclear molecules of the acetate-type seems to decrease in the following order: CH_3- $>Cl\cdot CH_2->Cl_2\cdot CH->Cl_3\cdot C-$ for a series of compounds with an empirical formula of $Cu(R\cdot COO)_2$, R being the afore-mentioned radical. This trend may be understood as due primarily to the accumulation of the inductive effect of the chlorine. Thus, as the hydrogen atom in the CH_3- group is replaced by the chlorine atom, the electron-withdrawing power of group R becomes greater to the effect that the electron density on the oxygen atom of the carboxylate group becomes smaller and consequently the electron on the copper atom would be withdrawn more strongly toward the oxygen atom. If it is assumed that the increase in the electronic density on the copper atom means increase of the electronic charge available for the additional linkage between the two copper atoms, the copper-to-copper linkage might become weaker as the electronic charge on the copper decreases on going from the trichloroacetate to the acetate.

Dimer Formation of Cupric Chloroacetates in Different Solvents.—The tendency to the dimer formation of cupric carboxylates in solution is found to depend greatly upon the solvent used. In the present work, water, ethanol, dioxane and ethyl ether as a solvent were compared with one another in respect of the dimer formation of the cupric carboxylates. Examinations were made about the structure of the cupric chloroacetates in these solvents in almost equal concentration. Examinations were also made about the structure of the component molecules in the crystals obtained after recrystallization from these solvents.

Absorption spectra of cupric mono-, di- and trichloroacetates in ethanol are shown in Fig. 6. From the above-mentioned criteria, it may be readily concluded that cupric monochloroacetate in ethanol exists mostly as the binuclear molecules, since the compound shows an absorption band at about $80\times10^{13}\,sec^{-1}$ in addition to a band at about $40\times10^{13}\,sec^{-1}$, the intensity of these bands being characteristic of

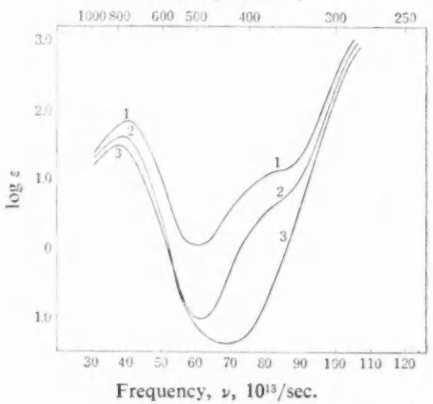


Fig. 6. Absorption spectra of $Cu(Cl_nH_{3-n}C\cdot COO)_2$ in ethanol: 1, $n=1$; 2, $n=2$; 3, $n=3$.

the binuclear cupric carboxylates. On the contrary, the dichloroacetate in ethanol shows only one band at about $40\times10^{13}\,sec^{-1}$ which is much weaker than the corresponding band of cupric monochloroacetate. Moreover, the compound shows no band at about $80\times10^{13}\,sec^{-1}$, although only a very small inflation was observed in this wavelength region. Judging from the criteria mentioned above, the dichloroacetate in ethanol exists mostly in the mononuclear form, although the binuclear molecule might be present in a very minor proportion. The trichloroacetate in ethanol is also seen to show an absorption spectrum, quite similar to the spectrum of the dichloroacetate in ethanol. It is, therefore, concluded that the trichloroacetate exists in the mononuclear form in this solvent. It is to be noted that the band at about $40\times10^{13}\,sec^{-1}$ of the monochloroacetate in ethanol seems to be a little weaker than the corresponding band of the same compound in dioxane, where it exists as binuclear molecules. Therefore, even the monochloroacetate in ethanol contains mononuclear complexes in a very small amount, which is in equilibrium with the binuclear molecules in most part.

In a similar way, the ultraviolet spectra indicate that cupric mono- and dichloroacetate in ethyl ether exist mostly as their binuclear molecules. It is to be noted that anhydrous cupric trichloroacetate shows an absorption spectrum in ethyl ether which is quite different from its spectrum in dioxan, ethanol or water. The spectrum of anhydrous trichloroacetate in ethyl ether was determined in the present work and is shown in Fig. 7. It is seen from the figure that anhydrous cupric trichloroacetate in ethyl ether shows the two absorption bands at about 38 and $80\times10^{13}\,sec^{-1}$, with the

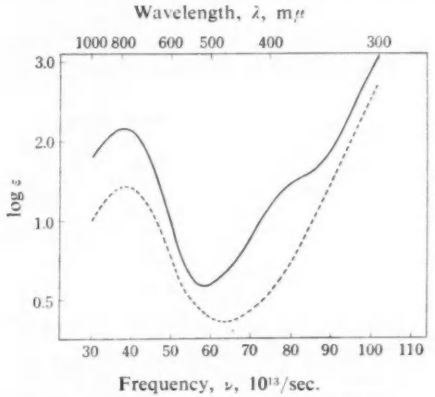


Fig. 7. Absorption spectra of anhydrous $Cu(Cl_3C\cdot COO)_2$ in ethyl ether (——) and in dioxane (----).

principal characteristics of the binuclear cupric carboxylates.

It is seen again that the band I of the cupric chloroacetates is displaced bathochromically, as the hydrogen atom is replaced by the chlorine atom in the series from the acetate to the trichloroacetate. This is in agreement with the relationship, which was derived in the case of dioxane as a solvent, and may be understood in a similar way. It is clearly seen that the cupric chloroacetates exist as binuclear molecules most readily in ethyl ether, of the solvents examined.

Thus the presentd work indicates that cupric trichloroacetate exists as mononuclear complexes in ethanol and in dioxane as well as in water, but that anhydrous cupric trichloroacetate exists in ethyl ether mostly as the binuclear molecules with the metal-to-metal linkage. Similarly it is found that cupric dichloroacetate exists as its binuclear form in dioxane and ether, but as the mononuclear form in ethanol and water. The monochloroacetate is found to exist mostly as binuclear molecules in ethyl ether, dioxane and ethanol, but as mononuclear molecules in water.

The dissociation of the dimer into the monomers takes place most readily in water. Even cupric acetate exists in this solvent mostly as monomeric complexes in equilibrium with the binuclear molecules in a small proportion, although the degree of dissociation depends upon the concentration of the solution. All the cupric monocarboxylates so far examined are found to exist in water as their monomers to a considerable extent if they are soluble in water at all.

Inspection of the data, which are summarized in Table II, reveals that the dimer formation of the cupric carboxylates in different solvents seems to become more difficult in the following order of the solvents: ethyl ether, dioxane, ethanol and water.

The order of the solvents, as presented above, is seen to agree with the increasing order of the electronic charge on the oxygen atom which also represents an increasing order of the co-ordinating capacity of the oxygen to

Table II. Structure of cupric chloroacetates in solution

	Water	Ethanol	Dioxane	Ethyl ether
Acetate	m	d	d	insoluble
Monochloro-〃	m	d	d	d
Dichloro- 〃	m	m	d	d
Trichloro- 〃	m	m	m	d

d: The compound exists mostly as binuclear molecules.
m: The compound exists mostly as mononuclear complexes.

the cupric ion. It is likely that the stronger co-ordination of the solvent molecules at the residual co-ordination position of the copper ion may break the binuclear molecules into the mononuclear complexes more easily. The present examinations also indicate that the mono- and the dichloroacetate give crystals consisting of the binuclear molecules, when recrystallized from water, ethanol or dioxane. On the contrary, the trichloroacetate gives from the same solvents no crystals of the binuclear molecules. The trichloroacetate, however, gives crystals of the binuclear molecules on recrystallization from ethyl ether after careful dehydration of the hydrated crystals. These observations, which are summarized in Table III, seem to be compatible with the above-mentioned order of the solvents concerning the dimer formation of the cupric monocarboxylates.

TABLE III. STRUCTURE OF THE COMPLEXES CONSTITUTING THE CRYSTALS OBTAINED BY RECRYSTALLIZATION OF CUPRIC CHLOROACETATES FROM VARIOUS SOLVENTS

	Water	Ethanol	Dioxane	Ethyl ether
Acetate	D	D	D	Insoluble
Monochloro-〃	D	D	D	D
Dichloro-〃	D	D	D	D
Trichloro-〃	M	M	M	D

D: Crystals consisting of dimeric molecules are obtained, aside from the question whether crystals of the monomeric complexes may be obtained or not.

M: Only crystals consisting of the monomeric complexes are obtained.

All these results are found to be in agreement with the conclusion that the trichloroacetate is the cupric chloroacetate which forms the binuclear molecule of the acetate-type with the greatest difficulty, the dichloroacetate being the next. Such a tendency may be understood as due primarily to the accumulation of the inductive effect of the chlorine, which seems to be in parallelism with the decrease in electronic density on the oxygen atom and with the decrease of dissociation constant of the free carboxylic acid. Among cupric carboxylates which are closely related with one another, the dissociation constant of the free carboxylic acid may be regarded as a measure of the formation of the binuclear molecule of the acetate-type.

Summary

In order to examine the structure of cupric chloroacetates, the dichroism in the visible and the ultraviolet region has been determined at room temperature by the microscopic method with mono-, di- and trichloroacetate. Visible and ultraviolet absorption spectra have also been determined with the above compounds in organic solvents.

The monochloroacetate consists of the binuclear molecules of the acetate-type in the crystalline state and in solution. It has been found that there is another form of the anhydrous monochloroacetate, which is colored blue and contains no binuclear molecule.

The dichloroacetate exists as binuclear molecules in ethyl ether and dioxane, but as mononuclear complexes in ethanol. The anhydrous dichloroacetate in the crystalline state consists of binuclear molecules of the acetate-type.

The trichloroacetate exists as mononuclear complexes in ethanol and in dioxane, and in the crystalline state. The compound exists as binuclear molecules in ethyl ether. Crystals of another form have also been obtained. These possibly consist of binuclear molecules.

The present authors wish to thank the Ministry of Education for a grant-in-aid, and Mr. H. Nakamura for experimental assistance at the earlier stage of this work.

Department of Chemistry
Faculty of Science
Osaka University
Nakanoshima Kita-ku, Osaka

The Reaction Products of 2,4,7-Tribromotropone and Anilines

By Tetsuo NOZOE, Kozo DOI* and Shigetaka ENDO**

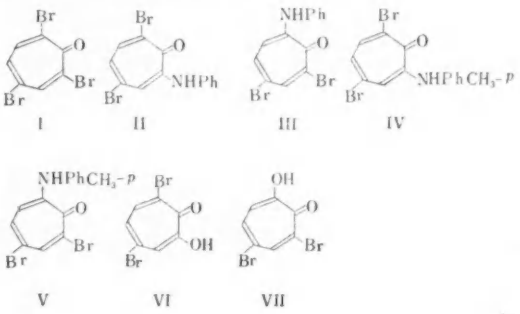
(Received March 11, 1960)

Nozoe and his co-workers[1] have revealed that the treatment of 2,4,7-tribromotropone (I) with aniline or *p*-toluidine affords a mixture of 2-anilino-4,7-dibromotropone (II) and 7-anilino-2,4-dibromotropone (III), or 4,7-dibromo-2-*p*-toluidinotropone (IV) and 2,4-dibromo-7-*p*-toluidinotropone (V); however, the respective components were not separated.

S. Iseda[2], in 1955, reported that the reaction of tribromotropone and aniline or *p*-toluidine resulted in the selective formation of 2-anilinotropone (II) or 2-*p*-toluidinotropone (IV), hydrolyses of which with ethanolic potassium hydroxide easily gave 3,6-dibromotropolone (VI).

The present authors have re-examined these reactions in order to prepare the dibromotropolone and have been able to isolate some reaction products besides the ones described above.

Heating of the tribromotropone (I) and aniline in benzene afforded 7-anilino-2,4-dibromotropone (III) in 21% yield, besides known 2-anilino-4,7-dibromotropone (II) in 66% yield. Separation of both products was easily effected by the application of the difference in their solubilities in chloroform. The structure of the former (III) was clarified through the formation of known 3,5-dibromotropolone (VII) on hydrolysis with alcoholic potassium hydroxide.



It had been found that the reaction of the tribromotropone and sodium acetate in acetic acid affords 3,6- (VI) and 3,5-dibromotropolones[1] (VII). In a similar manner, the above reaction of the tribromotropone and aniline was proved to give not only a 2-substituted but also a 7-substituted product, unlike that observed by Iseda[2].

The reaction of the tribromotropone and *p*-toluidine appeared to result in the formation of more complicated reaction products. Heating of the two compounds in benzene under the same condition as the above gave the 2-*p*-toluidinotropone (IV) in 55% yield and the 7-*p*-toluidinotropone (V) too in 5% yield, as was expected from the reaction in the case of aniline. The structure of the latter V was proved by the formation of known 3,5-dibromotropolone (VII) on its hydrolysis with alcoholic potassium hydroxide.

The ultraviolet absorption spectra of the above four kinds of anilinotropone are shown in Figs. 1 and 2. In general, the 7-anilinotropones (III and V) have absorption maxima in the longer wavelength region than the 2-anilinotropones (II and IV); this fact appears

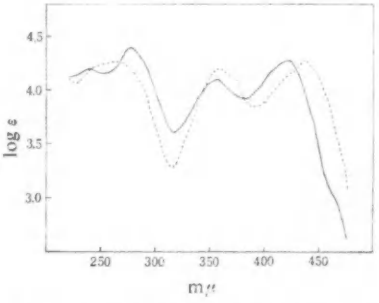


Fig. 1. Ultraviolet absorption spectra.
(——): 2-Anilino-4,7-dibromotropone (II)
(------): 7-Anilino-2,4-dibromotropone (III)

to be probably due to a stronger resonance effect of the bromine atom in the 4-position of the former (III and V) than that in the same position of the latter (II and IV) on the respective mesomeric interaction between the tropone ring and the benzene ring involving a nitrogen atom.

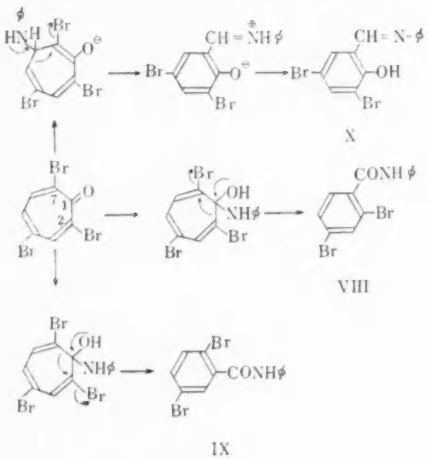
* Present address: The Chemical Research Institute of Non-Aqueous Solutions, Tohoku University, Katahira-cho, Sendai.
** Present address: Tokyo Research Laboratory, Dai-Nippon Celluloid Co., Ltd., Itabashi-ku, Tokyo.

1) T. Nozoe, Y. Kitahara, T. Ando, S. Masamune and H. Abe, *Sci. Repts. Tohoku Univ., Ser. I*, **36**, 166 (1952).
2) S. Iseda, This Bulletin, **28**, 617 (1955).

The reaction mixture, from which the foregoing toluidinotropones (IV and V) were removed, afforded furthermore four kinds of reaction product not isolated to date: 2,4-dibromobenz-*p*-toluidide (VIII), 2,5-dibromobenz-*p*-toluidide (IX), 3,5-dibromosalicylidene-*p*-toluidine (X), and an unidentified substance XI of m.p. 158°C.

The toluidide (VIII and IX) was easily converted into the known dibromobenzoic acid, respectively, on treatment with 75% sulfuric acid. The azomethine compound X was identified with a condensation product of 3,5-dibromosalicylaldehyde and *p*-toluidine. Another reaction product XI has the composition of $C_{14}H_{12}O_2NBr$, the structure of which will be elucidated in future.

It has been found that the tribromotropone (I) converts into benzoic acid derivatives by the action of bases such as dilute alkali or ammonia[1]. Formations of the dibromobenz-*p*-toluidides (VIII and IX) from the tribromotropone may be taken as the products of the benzilic acid type of rearrangement due to a nucleophilic attack of *p*-toluidine on the 1-position of the molecule. Formation of a Schiff's base X recalls that of 3-bromosalicylaldehyde from 2,7-dibromotropone by the action of alkali[3,4], and the mechanisms of the above reaction can be illustrated as follows:



X

VIII

IX

Experimental

The ultraviolet absorption spectra were measured in methanol with a Beckman model DU spectrophotometer.

Reaction of 2,4,7-Tribromotropone (I) and Aniline.—A mixture of the tribromotropone (I, 70 g.), anhydrous sodium acetate (21 g.), aniline (28 g.) and benzene (1400 ml.) was heated under reflux for 3 hr. The residue obtained by the evaporation of the solvent was washed with water and subsequently extracted with chloroform (100 ml). Crystallization of the part less soluble in chloroform gave 2-anilino-4,7-dibromotropone (II) as yellow needles from benzene, m.p. 197~198°C (reported[2] m.p. 196~197°C) (47.4 g., 66%). The part more soluble in chloroform, after evaporation of the solvent followed by crystallization from benzene, afforded 7-anilino-2,4-dibromotropone (III) as yellow needles, m.p. 179.5°C (15 g., 21%).

Found: C, 43.52; H, 2.46; N, 4.10. Calcd. for $C_{13}H_9ONBr_2$: C, 43.96; H, 2.56; N, 3.95%. λ_{max}^{MeOH} mμ (log ε): 265 (4.25), 361 (4.18), 438 (4.27).

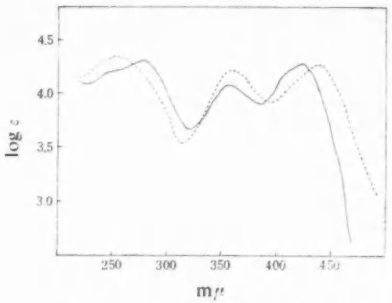


Fig. 2. Ultraviolet absorption spectra.
(——): 4,7-Dibromo-2-*p*-toluidinotropone (IV)
(------): 2,4-Dibromo-7-*p*-toluidinotropone (V)

Hydrolysis of 7-Anilino-2,4-dibromotropone (III).—A mixture of III (280 mg.), ethanol (10 ml.), potassium hydroxide (440 mg.) and water (4 ml.) was heated under reflux for 20 min. It was then diluted with water, acidified with concentrated hydrochloric acid to Congo-red and extracted with chloroform. The extract, after evaporation of the solvent followed by crystallization from methanol, gave 3,5-dibromotropolone (VII), m.p. 154~155°C, undepressed on admixture with an authentic specimen[5].

Reaction of Tribromotropone (I) and *p*-Toluidine.—A mixture of the tribromotropone (40 g.), anhydrous sodium acetate (11.6 g.), *p*-toluidine (13 g.) and benzene (700 ml.) was heated under reflux for 3 hr. The residue obtained by evaporation of the solvent was washed with water and extracted with acetone to separate into the less soluble A and the more soluble B portion. The fraction A, 37.3 g., after two recrystallizations from benzene, afforded 4,7-dibromo-2-*p*-toluidinotropone (IV) as yellow needles, m.p. 194.5°C (reported[2] m.p. 196°C) (24 g., 55%). The benzene mother-liquor, after chromatographic purification (benzene—alumina) followed by two fractional recrystallizations from the same solvent, gave a more soluble product, 2,4-dibromo-7-*p*-toluidinotropone (V) as yellow needles, m.p. 177°C (2.4 g., 5.6%).

Found: C, 45.68; H, 3.08; N, 3.67. Calcd. for

3) S. Seto, *Sci. Repts. Tohoku Univ., Ser. I*, **37**, 377 (1953).
4) Y. Kitahara, *ibid.*, **39**, 250 (1956).
5) T. Nozoe, Y. Kitahara, K. Doi and T. Arai, *Bull. Chem. Research Inst. Non-Aqueous Solns., Tohoku Univ.*, **7**, 13 (1957).

$C_{14}H_{11}ONBr_2$: C, 45.54; H, 3.01; N, 3.79%. λ_{max}^{MeOH} mμ (log ε): 250 (4.35), 361 (4.22), 440 (4.28).

From the same source a less soluble product XI was obtained as yellow plates, m. p. 157.5~158°C (2.3 g).

Found: C, 55.38; H, 4.02; N, 4.58. Calcd. for $C_{14}H_{12}O_2NBr$: C, 54.92; H, 3.95; N, 4.58%. λ_{max}^{MeOH} mμ (log ε): 240 (4.17), 283 (4.37), 345 (3.99), 350 (4.01), 400 (4.20).

The fraction B, after evaporation of the solvent, crystallization from ethyl acetate and chromatographic purification (benzene—alumina) followed by several recrystallizations from benzene, yielded the following three products in the order of decreasing of solubilities in the solvent:

2,5-Dibromobenz-*p*-toluidide (IX), colorless needles, m. p. 155~155.5°C (0.25 g.),

Found: C, 45.64; H, 2.85; N, 3.80. Calcd. for $C_{14}H_{11}ONBr_2$: C, 45.54; H, 3.01; N, 3.79%.

2,4-Dibromobenz-*p*-toluidide (VIII), colorless needles, m. p. 174°C (0.2 g.),

Found: C, 45.27; H, 2.70; N, 3.89%.

3,5-Dibromosalicylidene-*p*-toluidine (X), orange needles, m. p. 129.5°C (0.44 g.).

Found: C, 45.58; H, 2.92; N, 4.00%.

Hydrolysis of 2,4-Dibromo-7-*p*-toluidinotropone (V).—A mixture of the tropone (V, 90 mg), ethanol (5 ml.), potassium hydroxide (200 mg.), and water (3 ml.) was heated under reflux for 30 min., then acidified with dilute hydrochloric acid to Congo-red, and extracted with chloroform. The extract gave 3,5-dibromotropolone (VII, 40 mg.), m. p. 154~155°C, undepressed on admixture with an authentic sample[5].

Hydrolysis of 2,4-Dibromobenz-*p*-toluidide (VIII).—A mixture of the toluidide (VIII, 70 mg.) and 75% sulfuric acid (5 ml.) was heated at 130°C for 30 min. The solid obtained by pouring the reaction mixture onto crushed ice gave 2,4-dibromobenzoic acid, m. p. 166~169°C, undepressed on admixture with an authentic specimen[6].

Hydrolysis of 2,5-Dibromobenz-*p*-toluidide (IX).—Hydrolysis of IX under the same condition as the above yielded 2,5-dibromobenzoic acid, m. p. 151~153°C, undepressed on admixture with an authentic sample[7].

3,5-Dibromosalicylidene-*p*-toluidine (X).—A mixture of 3,5-dibromosalicylaldehyde[8] (30 mg.), *p*-toluidine (20 mg.) and ethanol (2 ml.) was heated under reflux for one hour. The reaction mixture, when cooled, afforded Schiff's base X as orange needles, m. p. 129~129.5°C, undepressed on admixture with the reaction product X of the tribromotropone and *p*-toluidine.

We gratefully acknowledge our debt to the Ministry of Education for the Grant in Aid for Fundamental Scientific Research and to Messrs. S. Ohyama, K. Kato, A. Iwanaga and S. Azumi for the microanalyses.

Department of Chemistry
Faculty of Science
Tohoku University
Katahira-cho, Sendai

6) A. K. Miller, *J. Chem. Soc.*, **61**, 1033 (1892).
7) H. Hubner, *Ber.*, **10**, 1705 (1877).
8) E. Tummeley, *Ann.*, **251**, 179 (1889).

Further Studies on the Intramolecular Cyclization of o,o'-Dihydroxydiphenyldiacetylene

By Fumio Toda and Masazumi Nakagawa

(Received February 27, 1960)

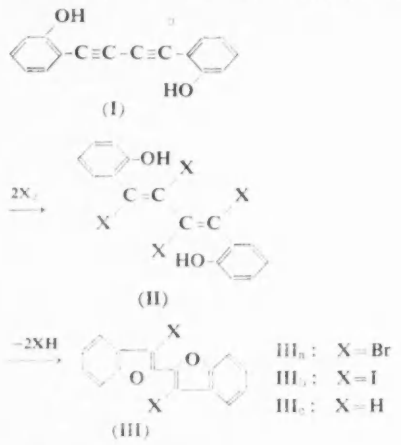
In the previous paper, the present authors reported the formation of 2,2'-dibenzofuranyl (III_c) from *o,o'*-dihydroxydiphenyldiacetylene (I) in alkaline madia[1]. The present paper deals with the same type of intramolecular cyclization of I and its silver salt IV in a neutral or an acidic medium.

Treatment of I in carbon tetrachloride with bromine yielded colorless needles in a 37% yield with evolution of hydrogen bromide. The structure of III_a was assigned to the substance from the analytical data, ultraviolet and infrared spectroscopic evidences (Figs. 1, 2 and Table I). The cyclization of I to III_a presumably involves the addition of bromine to the triple bands followed by the elimination of hydrogen bromide as indicated in the following scheme. The formation of diindole from *o,o'*-dinitrotolan which was reported by Ruggli and his associate seems to involve an essentially similar reaction[2]. In the case of the reaction of iodine with I under the same

1) F. Toda and M. Nakagawa, This Bulletin, **32**, 514 (1959).

2) P. Ruggli, *Ber.*, **50**, 883 (1917); P. Ruggli and O. Schmid, *Helv. Chim. Acta*, **18**, 1215 (1935).

condition, compounds III_b and III_c were isolated in a 10% and a 29% yield respectively.



When the reaction was carried out in the presence of finely powdered calcium carbonate in order to remove the evolved hydrogen iodide, the only product was found to be III_b. Treatment of III_b with hydroiodic acid resulted in the formation of III_c, therefore, the formation of III_c in the reaction of I with iodine is attributable to the reduction of III_b by

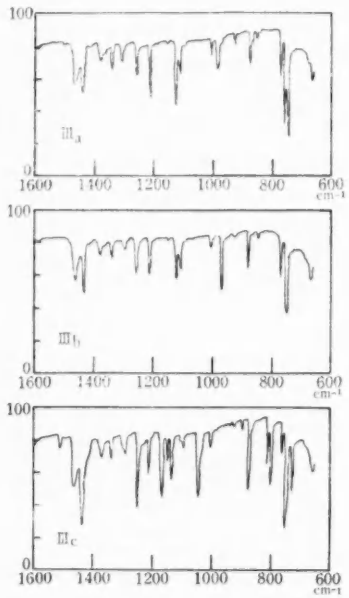


Fig. 1. The infrared spectra of 2,2'-dibenzofuranyl derivatives (Nujol mull)
III_a: 3,3'-Dibromo-2,2'-dibenzofuranyl
III_b: 3,3'-Diiodo-2,2'-dibenzofuranyl
III_c: 2,2'-Dibenzofuranyl

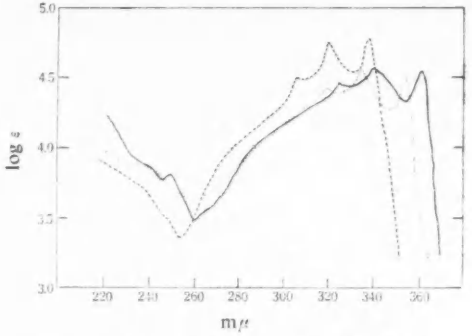


Fig. 2. The ultraviolet spectra of 2,2'-dibenzofuranyl derivatives.
·····: 3,3'-Dibromo-2,2'-dibenzofuranyl (III_a)
——: 3,3'-Diiodo-2,2'-dibenzofuranyl (III_b)
– – –: 2,2'-Dibenzofuranyl (III_c)

hydrogen iodide. The infrared and ultraviolet spectrum of III_b are shown in Figs. 1, 2 and Table I together with its mother substance and analogues. As illustrated in Fig. 1, the infrared spectra of III_a and III_b lack absorption at 800 and 812 cm^{-1} which exsists in the spectra of III_c and related 2-substituted benzofurans. Therefore, the absorption at ca. 800 cm^{-1} may be ascribed to the out-of-plane deformation of the hydrogen atom at the 3-position of the benzofuran ring.

A quantitative formation of III_c was observed when the silver salt IV was exposed to light for a few days. The reaction of methyl iodide with IV which was kept in a brown desiccator for a day gave 2-(*o*-methoxyphenylethynyl)-benzofuran (VI) together with III_c. The structure of VI was inferred from the close resemblance of its ultraviolet and infrared spectrum with those of 2-(*o*-hydroxyphenylethynyl)-benzofuran (VII)[1] as shown in Fig. 3 and Table I.

It is the most interesting result that the silver salt IV was converted into III_c and

TABLE I. ULTRAVIOLET SPECTRA OF 3,3'-DISUBSTITUTED 2,2'-DIBENZOFURANYL AND RELATED COMPOUNDS

Compound	Absorption (λ_{max} in mμ in 95% ethanol)				
III_a			319 (253)	335 (352)	354 (299)
III_b	250 (63)		325 (276)	342 (395)	361 (338)
III_c		306 (316)	320 (589)	338 (601)	
VI	250 (75)	299 (251)	318 (305)	340 (294)	
VII	250 (79)	298 (265)	319 (297)	340 (284)	

The figures in parentheses are $\varepsilon_{max} \times 10^{-3}$.

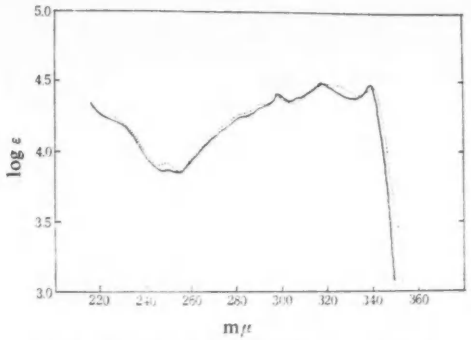


Fig. 3. The ultraviolet spectra of 2-substituted benzofurane derivatives.
——: 2-(*o*-Methoxyphenylethynyl)-benzofuran (VI).
·····: 2-(*o*-Hydroxyphenylethynyl)-benzofuran (VII).

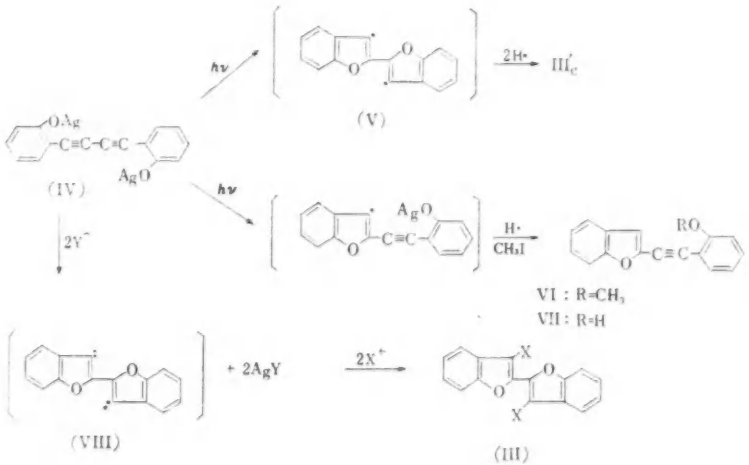
its derivatives in good yield by the reaction with various kind of reagents X–Y, such as hydrogen chloride, hydrogen iodide, acetic acid, potassium iodide, ammonium bromide, bromine, iodine and iodine chloride. The results are summarized in Table II.

The photochemical transformation of IV into III_c proceeds probably through a radical mechanism. The photodecomposition of IV may produce a phenoxy radical, and the intramolecular addition of the oxygen radical to the triple bonds would yield an intermediate dibenzofuranyl radical (V) as illustrated in the following scheme. The radical may be stabilized

Table II. Reaction of the silver salt IV with X–Y

Reagent X	Reagent Y	Medium	Product
H	Cl	Water	2,2'-Dibenzofuranyl (III_c)
H	I	〃	〃
H	CH_3CO_2	Acetic acid	〃
K	I	Water	〃
NH_4	Br	〃	〃
Br	Br	〃	3,3'-Dibromo-2,2'-dibenzofuranyl(III_a)
I	I	Carbon tetrachloride	3,3'-Diiodo-2,2'-dibenzofuranyl(III_b)
I	Cl	Water	〃

by the substruction of hydrogen atoms from water or ethanol molecule yielding III_c*. Also the formation of VI is quite understandable assuming that the light induced half ring closure of IV is followed or is preceded by the ionic displacement of the silver atom by the methyl group. A same type of photochemical reaction of an acetylenic compound has been reported by Prey[3]. He found that the silver acetylide of *o*-hydroxyphenylacetylene yields benzofuran when irradiated with a quarz lamp. On the other hand, the above mentioned reactions of IV with various kinds of argentphilic reagents presumably proceed through an ionic mechanism. The substruction of silver cation from IV by the attack of the anionic moiety of the reagent Y^- should result in the formation of 2,2'-dibenzofuranyl anion (VIII). The combination of the cationic moiety of the reagent X^+ or proton of the



3) V. Prey and G. Pieh, *Monatsh.*, **80**, 790 (1949).
* The silver salt was prepared as a pasty mass by the addition of an aqueous silver nitrate to a solution of I in ammoniacal ethanol.

solvent with VIII gives dibenzofuranyl (III_c) or its 3,3'-disubstituted derivatives.

The bathochromic shifts of the spectrum of III_a, III_b and III_c are found to be in the expected sequence of H < Br < I with respect to the substituents at the position of 3 and 3'. On the other hand, the hyperchromic effect of the substituents is in the sequence of H < I < Br. The low ε-values of the K-band of III_a and III_b as compared with that of III_c may be attributable to an increase of twist of the molecule around the 2,2'-single bond caused by the presence of bulky halogen atoms at 3- and 3'-position. Similar phenomena have already been observed in several instances[4], and were explained in terms of the diminished probability of excitation of the molecule to a coplanar excitation state.

Experimental*

Formation of 3,3'-Dibromo-2,2'-dibenzofuranyl (III_a) from *o,o'*-Dihydroxydiphenyldiacetylene (I).—Bromine (3.2 g.) was added to a suspension of finely powdered I (2.4 g.) in dry carbon tetrachloride (40 cc.). An evolution of hydrogen bromide under an exothermic reaction was observed. The mixture was shaken for one hour with occasional cooling with water. The crystals separated were filtered, and the solvent was evaporated. The remaining solid was combined with the crystals, washed with hot ethanol. Recrystallization from benzene yielded III_a as colorless needles, 1.5 g. (37%), m. p. 225°C.

Found: C, 48.58; H, 2.03; Br, 41.02. Calcd. for $C_{16}H_8O_2Br_2$: C, 49.02; H, 2.05; Br, 40.77%.

I. R. max., 1255 (=C-O-), 875(furan ring)cm^{-1}.

Formation of 3,3'-Diiodo-2,2'-dibenzofuranyl (III_b) and 2,2'-Dibenzofuranyl (III_c) from *o,o'*-Dihydroxydiphenyldiacetylene (I).—A mixture of I (1.2 g.), dry carbon tetrachloride (25 cc.) and iodine (2.6 g.) was allowed to stand for 3 days at room temperature. The solvent was removed by distillation, and the careful fractional recrystallization of the residual solid from ethanol resulted in the separation of two substances. The less soluble substance was found to be III_b, 0.25 g. (10%), m. p. 212°C.

Found: C, 39.59; H, 1.59; I, 52.58. Calcd. for $C_{16}H_8O_2I_2$: C, 39.54; H, 1.64; I, 52.22%.

I. R. max., 1252 (=C-O-), 875 (furan ring)cm^{-1}.

The more soluble substance, 0.35 g. (29%), m. p. 193~194.5°C was identified as III_c by a mixed melting point determination with an authentic specimen.

A mixture of I (0.95 g.), dry carbon tetrachloride (30 cc.), iodine (2.06 g.) and finely powdered calcium carbonate (1.6 g.) was kept at room temperature for 2 days. The inorganic salt was removed by filtration and extracted with hot benzene. The extract was combined with the filtrate and the solvent was removed under reduced pressure. The residue thus obtained was digested with hot benzene. The crude crystals obtained by the evaporation of the solvent in vacuo was recrystallized from ethanol yielding 0.32 g.. (16%) of III_b, m. p. 212°C as colorless needles.

Reduction of III_b with Hydrogen Iodide.—Hydroiodic acid (57% aqueous solution, 4 g.) was added to a solution of III_b (50 mg.) in carbon tetrachloride (30 cc.), and the mixture was shaken for 2 days at room temperature. Treatment of the reaction mixture afforded crystalline solid. Fractional recrystallization of the crude solid gave recovered III_a, 30 mg., m. p., 211~212°C and a small amount of III, m. p. 190~194°C.

Preparation of the Silver Salt of *o,o'*-Dihydroxydiphenyldiacetylene (IV).—Silver nitrate (2.8 g.) in water (5.0 cc.) was added to the solution of I (1.8 g.) in a mixture of concentrated ammonia (1.2 cc.) and ethanol (15 cc.). The light yellow precipitate formed was collected by filtration and washed with ethanol, yielding 2.8 g. (ca. 80%) of the silver salt (IV). The attempt to determine the melting point of the saltfailed. A gradual decomposition was observed, when the salt (IV) was heated in a capillary tube, and the substance in the tube fused at 194.5~195.5C showing the melting point of III_c.

Photodecomposition of the Silver Salt (IV).—Finely powdered IV was exposed to a diffused light in the laboratory for a few days. The resulted greyish powder was digested with benzene. The solvent was removed by distillation, and the residue was recrystallized from benzene or ethanol yielding colorless needles, m. p. 194.5~195.5°C in an almost quantitative yield. This substance was identified as III_c by a mixed melting point determination with an authentic specimen.

Formation of 2,2'-Dibenzofuranyl (III_c) from the Silver Salt IV.—A mixture of freshly prepared IV and aqueous hydroiodic acid (57%) was shaken for 5 min. The precipitate formed was collected and washed with water and ethanol, successively. The residue was digested with hot benzene. Concentration of the benzene solution afforded crude crystals. Recrystallization of the crude product from ethanol resulted in colorless needles, m. p. 194.5~195.5°C in good yield. The mixed melting point of this substance with authentic III_c showed no depression. Similar treatment of IV with concentrated hydrochloric acid, acetic acid, aqueous solution of potassium iodide (30%) and aqueous solution of ammonium bromide (30%) also yielded III_c in good yield.

Formation of 3,3'-Dibromo-2,2'-dibenzofuranyl (III_a) from the Silver Salt IV.—To a suspension of a freshly prepared IV in water, bromine was added drop by drop with shaking until the color of bromine persists. The insoluble material was collected by filtration, and washed with water and ethanol successively. The residue was digested with hot benzene and the solvent was evaporated. The crude crystals thus obtained were recrystallized from ethanol yielding colorless needles, m. p. 220~223°C, which showed no depression of the melting point on admixture with previously obtained III_a.

4) E. A. Braude, F. Sondheimer and W. F. Forbes, *Nature*, 173, 117 (1954); F. Toda and M. Nakagawa, This Bulletin, 33, 223 (1960).

* No melting points are corrected.

Formation of 3,3'-Diiodo-2,2'-dibenzofuranyl (III_b) from the Silver Salt IV.—A mixture of freshly prepared IV, iodine and dry carbon tetrachloride was shaken for 5 min. Silver iodide formed was removed by filtration. The crude crystals obtained by the evaporation of the solvent were recrystallized from ethanol resulting in colorless needles, m. p. 211~212°C. This substance showed no depression of the melting point on admixture with authentic III_b.

The reaction of iodine chloride with IV under the same condition also gave III_b in good yield.

Formation of 2-(*o*-Methoxyphenylethynyl)-benzofuran (VI) from the Silver Salt IV.—The silver (IV, 2.8 g.) which was kept in a brown desiccator for a day was mixed with methyl iodide (1.5 g.) and dry benzene (5.0 cc.). The mixture was refluxed for 1.5 hr. The precipitate formed was collected by filtration, and extracted with hot benzene (30 cc.). The crude crystals obtained by the evaporation of the solvent was digested with cold ethanol (15 cc.). The insoluble material was recrystallized from benzene resulting in colorless needles, 1.08 g. (78%), m. p. 194.5~195.5°C. The mixed melting point of this substance with 2,2'-dibenzofuranyl (III_c) showed no depression. The alcoholic solution was diluted with water, and the crude crystals formed were recrystallized from a diluted ethanol resulting in VI, colorless needles, 100 mg. (6%), m. p. 72°C.

Found: C, 82.33; H, 4.94. Calcd. for $C_{17}H_{12}O_2$: C, 82.24; H, 4.87%.

I. R. max., 1246, 1265, 1040 (=C-O-), 2200 (-C≡C-), 884 (furan ring), 807, 792 (δ CH of 3-position of furan ring) cm^{-1}.

The present authors wish to thank the Ministry of Education for a grant-in-aid. The authors are also grateful to Dr. T. Shiba, Mr. M. Okumiya and Miss K. Koike for the microanalyses.

Department of Chemistry
Faculty of Science
Osaka University
Nakanoshima, Osaka

Cyclic Acetylenes. III. Synthesis of 1,1'-Dianthryldiacetylene. A Model Substance of a Cyclic Tetraacetylene Containing Anthracene Nuclei*

By Shuzo AKIYAMA and Masazumi NAKAGAWA

(Received March 24, 1960)

In the course of the synthetical studies on the cyclic tetraacetylene containing two anthracene nuclei (I)[1], the authors have studied the synthesis of 1,1'-dianthryldiacetylene (X) as a model substance of I.

(I)

After several experimentations, the following route to X was developed. 1-Carboxyanthraquinone (III)[2] which was derived from 1,9-benz-10-anthrone (II) by chromic acid oxidation was converted to 1-carboxyanthracene (IV)[3] by reduction with zinc dust in aqueous ammonia. The acid chloride V[4] which was obtained on treatment of IV with thionyl chloride was caused to react with diethyl ethoxymagnesiomalonate in toluene. The ketomalonate VI thus obtained was treated without further purification with a mixture of propionic acid and sulfuric acid to cause ester exchange and decarboxylation. The over-all yield of 1-acetylanthracene (VII)[5] from II was found to be 52%. The attempt to prepare VII by the reaction of dimethyl cadmium with V was found to be unsatisfactory.

The ketone VII was treated with phosphorus pentachloride in phosphorus oxychloride, and the crude chloride VIII was dehydrochlorinated with sodium amide in liquid ammonia yielding 1-ethynylanthracene (IX) as light yellow cubes. This substance decomposed gradually on standing at room temperature. The presence of a terminal acetylene was confirmed by the formation of yellow cuprous

* For Part II, see F. Toda and M. Nakagawa, This Bulletin, 33, 230 (1960).

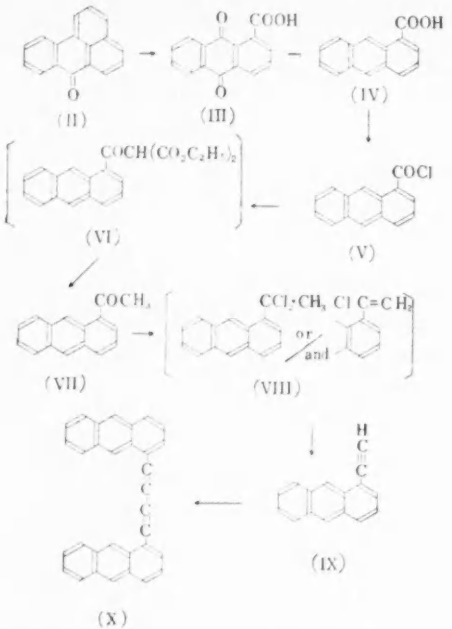
1) The following paper.

2) E. B. Barnett and J. W. Cook, *Ber.*, **57**, 1777 (1924).

3) C. Graebe and S. Blumenfeld, ibid., **30**, 1118 (1897).

4) J. W. Cook, *J. Chem. Soc.*, **1931**, 566; R. Scholl and J. Donat, *Ber.*, **62**, 1297 (1929).

5) VII can be obtained by the fractional recrystallization of a mixture with the 2-isomer which is the product of the direct acetylation of anthracene by the Friedel-Crafts reaction. The authors are grateful to Professor P. H. Gore for private communication suggesting the details of the above mentioned process. cf., P. H. Gore, *J. Org. Chem.*, **22**, 135 (1957).



and silver acetylides. Also the infrared spectrum of IX showed an absorption at 3270 cm^{-1} indicating the presence of the ethynyl group as illustrated in Fig. 1.

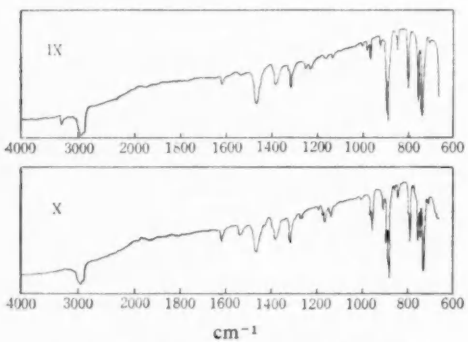


Fig. 1. The infrared spectra of 1-Ethynylanthracene (IX) and 1,1'-dianthryldiacetylene (X). (Nujol mull)

The oxidative coupling of IX according to the procedure of Eglinton[6] afforded fine yellow needles in almost quantitative yield. This was identified as X from its analytical data and the absence of free ethynyl absorption in the infrared spectrum (Fig. 1). X was found to be a stable substance and could be kept unchanged for a long time. X is slightly soluble in benzene and gives a greenish yellow solution with intense blue fluorescence. The diyne X is almost insoluble in common organic solvents such as ethanol and petroleum benzine.

The ultraviolet spectra of IX and X will be discussed in the following paper comparing them with those of the analogous compounds.

6) G. Eglinton and A. R. Galbraith, *Chem. & Ind.*, **1956**, 737; *J. Chem. Soc.*, **1959**, 889.

Experimental**

1-Carboxyanthraquinone (III).—The acid (III) was prepared by the chromic acid oxidation of 1,9-benz-10-anthrone (II) according to the method of J. W. Cook[2].

1-Carboxyanthracene (IV).—III was reduced by means of zinc dust and aqueous ammonia[3], m. p. 245°C, yellow rods (the reported value, m.p. 245°C[3]).

Found: C, 80.83; H, 4.63. Calcd. for $C_{15}H_{10}O_2$: C, 81.06; H, 4.54%.

1-Acetylanthracene (VII).—IV was converted to the corresponding acid chloride V by the reaction with thionyl chloride. Diethyl ethoxymagnesiomalonate was prepared from diethyl malonate (48 g., 0.3 mol.), magnesium (7.3 g., 0.3 gram atom) and absolute ethanol (27.6 g., 0.6 mol.) in dry toluene (150 ml.). The excess of ethanol was removed by the azeotropic distillation with toluene. To the refluxed solution of the malonate a solution of crude V (24 g., 0.1 mol.) in dry toluene (400 ml.) was added in 30 min., and the stirring was continued for 3 hr. also under reflux. The reaction mixture was chilled with ice, and decomposed by the addition of 6 N sulfuric acid (100 ml.). The organic layer separated was washed thoroughly with water and dried over anhydrous sodium sulfate. The solvent was removed under reduced pressure. The crude ketomalonate VI was obtained as an orange-yellow oil. The crude product was mixed with propionic acid (80 ml.) and concentrated sulfuric acid (0.5 ml.), and the mixture was refluxed. After 3 hr., 4 N sulfuric acid (50 ml.) was added and the refluxing was further continued for 8 hr. to complete decarboxylation. A small amount of toluene was co-distilled out with water from the reaction mixture. The residue was diluted with water (300 ml.), and the mixture was extracted with benzene (100 ml.×6). The extract was washed with dilute sodium hydroxide and water, successively. The dried benzene solution was concentrated to 200 ml. and filtered through a short column of alumina (20 g.). Light yellow cubes, (15 g., 68%), m. p. 103~106°C were obtained on concentrating the filtrate. Recrystallization of the crude crystals from benzene afforded pure VII as light yellow cubes, m. p. 106~107°C (the reported value, m. p. 105~106°C[3]).

Found: C, 87.52; H, 5.50. Calcd. for $C_{16}H_{12}O$: C, 87.24; H, 5.49%.

I. R. max., 1670 cm^{-1} (C=O).

1-Ethynylanthracene (IX).—The acetyl compound VII (10 g., 0.046 mol.) and phosphorus pentachloride (9.8 g., 0.068 mol.) was added to

** All melting points are uncorrected. The infrared spectra were measured by the Nujol mull method.

phosphorus oxychloride (70 ml.) and the mixture was refluxed for 1 hr. The mixture changed to a homogeneous dark red solution with an evolution of hydrogen chloride. Phosphorus oxychloride was removed completely under reduced pressure and the residual dark brown viscous oil was kept for several hours in an evacuated desiccator containing phosphorus pentoxide and potassium hydroxide.

The solution of the crude chloride in dry tetrahydrofuran (20 ml.) was added in 1 hr. to a solution of sodium amide (prepared from 11.5 g. of sodium) in liquid ammonia (500 ml.). After evaporation of ammonia, water (100 ml.) was added to the reaction product and extracted with ligh petroleum (b. p. 60~80°C, 1000 ml.). The extract was washed with water and dried over anhydrous sodium sulfate. The dried solution was filtered through a thin layer of alumina (40 g.). Concentration of the filtrate resulted in IX as light yellow cubes, 5.1 g. (55%), m. p. 88~90°C. This material was recrystallized from light petroleum or sublimed under high vacuum (0.5×10^{-4} mmHg) to yield pure IX, m. p. 90°C.

Found: C, 95.04; H, 5.10. Calcd. for $C_{16}H_{10}$: C, 95.02; H, 4.98%.

I. R. max., 3270cm^{-1} (-C≡C-H).

U. V. max. (in cyclohexane), λ_{max} 255 mμ, (log ε =5.05), 274.5 (4.04), 305 (2.60), 322 (3.03), 333.5 (3.44), 352 (3.78), 371 (4.04), 386 (3.68)* and 391 (4.06). Asterisk indicates shoulder.

Oxidative Coupling of 1-Ethynylanthracene (IX).—Cupric acetate monohydrate (40 g.) was added to the mixture of IX (1.89 g., 0.0094 mol.), dry pyridine (300 ml.) and methanol (20 ml.), and the mixture was stirred for 9 hr. at 55°C. The reaction mixture was filtered and the solid obtained was washed thoroughly with water. Yellow powder almost free from inorganic material was obtained, 1.45 g. (85%), m. p. 287~290°C (decomp.). The residue which was obtained by the concentration of the filtrate under a reduced pressure was digested with benzene. The benzene solution was passed through a short column of alumina and the filtrate was evaporated to yield yellow powder, 0.17 g. (10%), m. p. 289~291°C (decomp.). This was combined with the first crop and recrystallized from benzene to give pure X, yellow fine needles, m. p. 289~291°C (decomp.).

Found: C, 95.61; H, 4.47. Calcd. for $C_{32}H_{18}$: C, 95.49; H, 4.51%.

X gave no precipitate with an alcoholic silver nitrate solution.

U. V. max. (in cyclohexane), λ_{max} 259 mμ (log ε =5.16), 283 (4.27), 295 (4.35), 315.5 (3.52), 329 (3.66), 346 (3.64)*, 365 (3.99), 386 (4.26), 407 (4.42) and 428 (4.39). Asterisk indicates shoulder.

The authors are indebted to Mr. M. Okumiya and Miss K. Koike for the microanalyses. A part of the expenses of this research was defrayed from the grant-in-aid of the Ministry of Education to which the authors' thanks are also due.

Department of Chemistry
Faculty of Science
Osaka University
Nakanoshima, Osaka

*Cyclic Acetylenes. IV. Synthesis of a Cyclic Tetraacetylene Containing Anthracene Nuclei**

By Shuzo AKIYAMA, Soichi MISUMI and Masazumi NAKAGAWA

(Received March 24, 1960)

There has already been discussion concerning the fact that a conjugated planar cyclic en-yne system may form a continuous molecular π-orbital overlapping the *p*-orbitals of the carbon atoms of the unsaturated linkages in the cycle[1]. It is of interest to synthesize such a cyclic acetylene in order to establish experimentally the validity of the theoretical prediction, but the only member of this series known up to now is the completely conjugated planar eighteen membered monocyclic hexaene-triyne which was reported recently by Sondheimer and Wolovsky[2]. They synthesized the hexaene-triyne by a prototropic rearrangement of the corresponding cyclic hexa acetylene which was prepared by the oxidative coupling of 1,5-hexadiyne. Another promising approach to the synthesis of a planar cyclic en-yne system is the oxidative coupling of a diethynyl aromatic compound in which the two ethynyl groups are directly attached to the proper positions of the aromatic nucleus. The rigid geometry of this type of molecule may be favorable to the ring formation by oxidative coupling as pointed out by W. Baker[3]. The

* For a preliminary report of a part of this work, see *Chem. & Ind.*, **1960**, 346.

1) T. J. Sworski, *J. Chem. Phys.*, **16**, 550 (1948).

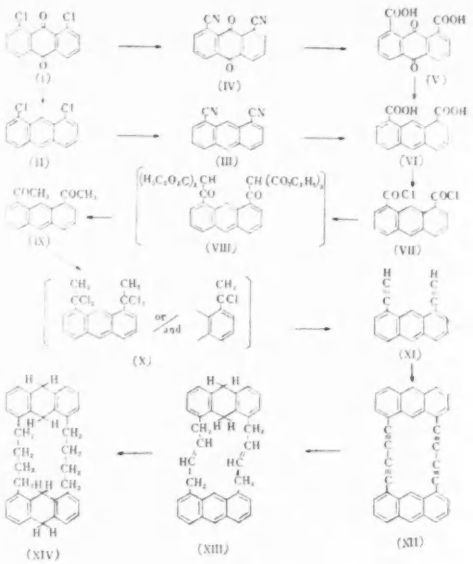
2) F. Sondheimer and R. Wolovsky, *J. Am. Chem. Soc.*, **81**, 1771 (1959).

first attempt according to this principle has been reported by Eglinton and his coworker. They carried out the oxidative coupling of *o*-diethynylbenzene expecting the formation of a fully conjugated strainless cyclic trimer, the tribenzo derivative of cyclo-octadeca-trienehexayne[4]. But unexpectedly the product was identified to be a highly strained cyclic dimer of *o*-diethynylbenzene[5].

We have synthesized a planar conjugated cyclic tetraacetylene containing two anthracene nuclei (XII) according to the following scheme. The starting material for our work was 1,8-dichloroanthraquinone (I). The dichloroquinone I was reduced to 1,8-dichloroanthracene (II)[6], and the reaction of II with cuprous cyanide in quinoline gave 1,8-dicyanoanthracene (III). The attempt to obtain 1,8-diacetylanthracene (IX) by the reaction of III with methylmagnesium iodide in anisol failed on account of the poor solubility and inert reactivity of III. Therefore, according to the results of the preliminary experiments[7], the route to the diacetyl compound IX via dicarboxylic acid VI was chosen. 1,8-Dicarboxyanthracene (VI) was synthesized by the alkaline hydrolysis of the dinitrile III in ethylene glycol. The diacid VI was also prepared according to the procedure of Waldmann[8] by the reduction of 1,8-dicarboxyanthraquinone (V) which was derived from I via IV.

Treatment of VI with thionyl chloride in chlorobenzene afforded the diacid dichloride VII. VII was caused to react with diethyl ethoxymagnesiomalonate in xylene to yield the ketomalonate VIII. The curde VIII was heated without further purification with the mixture of propionic and sulfuric acid. Treatment of the reaction mixture resulted in 1,8-diacetylanthracene (IX) as light yellow cubes.

The crude chloride X which was obtained by the reaction of the diacetyl compound IX with phosphorus pentachloride in phosphorus oxychloride was dehydrochlorinated with sodium amide in liquid ammonia. 1,8-Diethynylanthracene (XI) was obtained as yellow cubic crystals which melted at 150~151°C with decomposition. XI was found to be a fairly unstable substance and could not be kept without de-composition for a long time. The infrared spectrum of XI shows an absorption at 3260 cm^{-1} indicating the presence of terminal acetylene as illustrated in Fig. 1. XI gives also red cuprous and yellow silver acetylides.



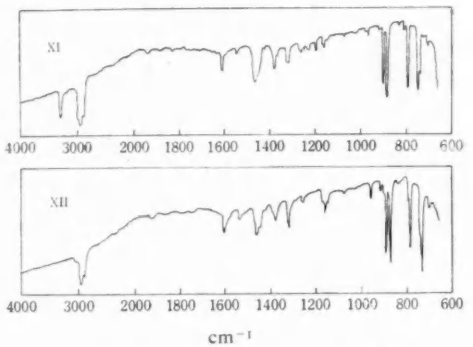


Fig. 1. The infrared spectra of 1,8-diethynylanthracene (XI) and the cyclic tetraacetylene (XII). (Nujol mull)

The oxidative coupling of the diethynyl compound XI, conducted according to the procedure of Eglinton[9], yielded orange crystals in a high yield. The structure of the cyclic tetraacetylene (XII) was assigned to the substance from the method of preparation and from the following evidences. The cyclic nature of XII was confirmed by the absence of free ethynyl absorption in the infrared spectrum (Fig. 1) and by the negative test against alcoholic silver nitrate solution. Various

3) W. Baker, J. F. W. McOmie and W. D. Ollis, *J. Chem. Soc.*, **1951**, 200.
4) G. Eglinton and A. R. Galbraith, *Proc. Chem. Soc.*, **1957**, 350.
5) O. M. Behr, G. Eglinton and R. A. Raphael, *Chem. & Ind.*, **1959**, 699; W. K. Grant and J. C. Speakmann, *Proc. Chem. Soc.*, **1959**, 231.
6) H. Schilling, *Ber.*, **46**, 1066 (1913).
7) The preceding paper.
8) a) H. Waldmann and A. Oblath, *Ber.*, **71**, 366 (1938). b) H. Waldmann and R. Stengl, *Chem. Ber.*, **83**, 167 (1950).
9) G. Eglinton and A. R. Galbraith, *Chem. & Ind.*, **1956**, 737; *J. Chem. Soc.*, **1959**, 886.

methods for determining the molecular weight of XII were tried unsuccessfully as a result of the poor solubility of this substance in organic solvents. Reduction of XII with an excess of sodium in liquid ammonia yielded a substance with a composition $C_{36}H_{30}$ as a main product.

The infrared spectroscopy of this material revealed the presene of a methylene group (ν_{as}: 2860, ν_s: 2920 cm^{-1}), unconjugated *trans*-ethylenic bond (δ_{CH}: 952 cm^{-1}) and 1, 2, 4, 5-tetrasubstituted benzene (δ_{CH}: 875 cm^{-1}) as illustrated in Fig. 1. From these evidences,

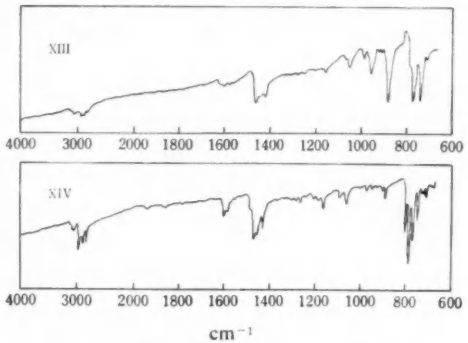


Fig. 2. The infrared spectra of $C_{36}H_{30}$ (XIII) and the hydrogenated product (XIV). (KBr-disk)

the structure of XIII was assigned tentatively to the reduction product. Further hydrogenation of XIII in benzene over palladium on charcoal gave tiny colorless cubes with the empirical formula $(C_9H_9)_n$. The molecular weight of this substance could be determined accurately by Rast's method in camphor. The molecular weight thus determined (450; Calcd. for $C_{36}H_{36}$: 469) provided an unequivocal evidence regarding the dimeric cyclic acetylene for XII. The fully hydrogenated product shows no absorption at 875 cm^{-1} in the infrared spectrum, indicating the absence of the anthracene nucleus in the molecule. The infrared spectrum of this substance was illustrated in Fig. 2 together with that of XIII. The ultraviolet spectrum of the hydrogenated product also indicates the destruction of anthracene chromophor. From these observations structure XIV was given to the reduction product.

The high yield of the oxidative cyclic dimerization of XI to XII seems to reflect the favorable orientation of the ethynyl groups in XI. Examination of the bond angles of the compound XII readily verifies its planar and strainless nature. XII was found to be a very stable compound and can be kept without change for a long time in air and light. In

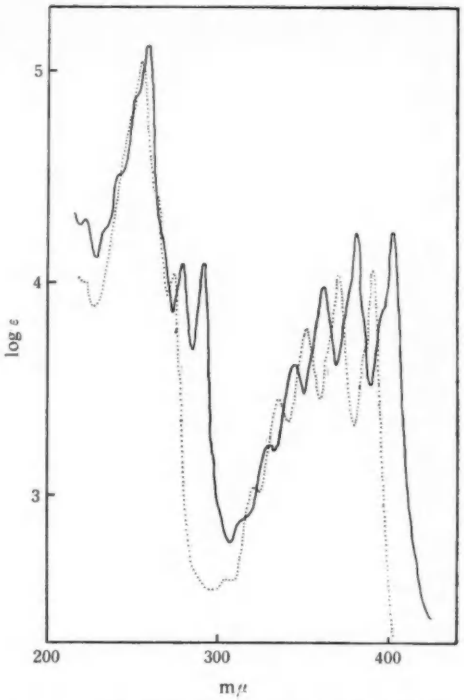


Fig. 3. The ultraviolet spectra of 1-ethynylanthracene (XVI, ·····) and 1,8-diethnylanthracene (XI, ——).

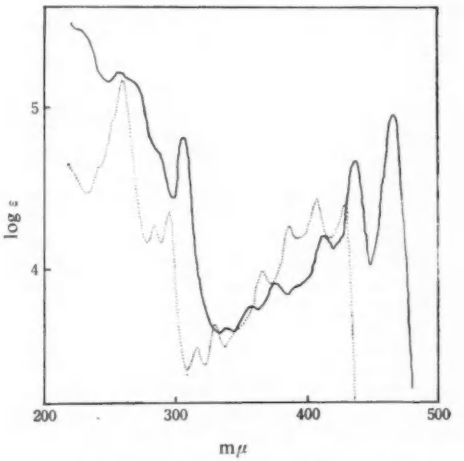


Fig. 4. The ultraviolet spectra of 1,1'-dianthryldiacetylene (XVIII, ·····) and the cyclic tetraacetylene (XII, ——).

an evacuated capillary tube, XII turned to black at a temperature over 370°C without the change of crystal form. XII is sparingly soluble in common organic solvents. The

Table I. The bathochromic shifts of the ultraviolet spectra of the acetylenes

Compound	λ_{max}	ε	Solvent	Shifts in mμ from XV	XVI	XI	XVII
Anthracene (XV)	374	7,460	C	—	—	—	—
1-Ethynylanthracene (XVI)	391	11,500	C	17	—	—	—
1,8-Diethynylanthracene (XI)	402	17,400	C	28	11	—	—
1,1'-Dianthryldiacetylene (XVII)	428	24,600	C	54	37	26	—
Cyclic tetraacetylene (XII)	465	91,200	B	91	74	63	37

C: cyclohexane, B: benzene.

solution in dimethylformamide, in which XII is slightly soluble, exhibits an orange yellow color with intense and characteristic green fluorescence.

The ultraviolet spectra of 1-ethynylanthracene (XVI)[7], 1,8-diethynylanthracene (XI), 1,1'-dianthryldiacetylene (XVII)[7] and the cyclic tetraacetylene (XII) are illustrated in Figs. 3 and 4. The bathochromic shifts of the absorption maxima at the longest wavelength and the ε-values of these acetylenes as compared with anthracene are tabulated in Table I. The ε-value of XVII reduced to per acetylene-unit ($0.5\times24,600=12,300$) is approximately the same as that of XVI (11,500), whereas the ε-value of XII reduced per diyne-unit ($0.5\times91,200=45,600$) is much greater than those of XVII (24,600) and XI (17,400). The marked increase in the absorption intensities and the remarkable red-shift of the spectrum of XII indicate that the cyclic tetraacetylene can not be regarded as a simple derivative of anthracene. These ultraviolet spectroscopic properties and, at least to some extent, the unusual stability of XII seem to be attributable to an enhanced interaction of π-electrons of the acetylenic bonds with those of the aromatic system.

The syntheses of the same type of higher cyclic acetylenes are now being investigated.

Experimental**

1,8-Dichloroanthracene (II).—The reduction of I with zinc dust in aqueous ammonia[6] yielded a product consisting of mainly 1,8-dichloro-9,10-dihydroanthranol[10]. The reduction product was extracted with ethanol. The solvent was removed and the light yellow crystalline residue was recrystallized from carbon tetrachloride to give the pure anthranol, white needles, m. p. 136~136.5°C.

Found: C, 63.52; H, 4.03; Cl, 27.27. Calcd. for $C_{14}H_{10}OCl_2$: C, 63.41; H, 3.78; Cl, 26.80%.

I. R. max., Broad absorption near 3200 (-OH); 1180 cm^{-1} (C-O).

A catalytic amount of concentrated sulfuric acid was added to a solution of the anthranol in hot ethanol. Immediate precipitation of practically pure II, light yellow needles, m. p. 157~158°C, (values in literature, 156°C[6]; 160°C[11]) was observed.

Found: C, 68.04; H, 3.28; Cl, 28.96. Calcd. for $C_{14}H_8Cl_2$: C, 68.05; H, 3.26; Cl, 28.69%.

In practice it was found more convenient to treat the ethanol extract of the reduction product with a small amount of sulfuric acid without the isolation of the anthranol.

1,8-Dicyanoanthracene (III).—A mixture of II (50 g., 0.19 mol.), cuprous cyanide (60 g.) and quinoline (500 ml.) was stirred for 24 hr. under reflux. The cooled reaction mixture was filtered and the solid was washed with dilute hydrochloric acid, then treated with dilute aqueous sodium cyanide to dissolve cuprous cyanide, and was washed with water. The dried light brown powder was digested with nitromethane (250 ml.) to remove the resinous material on the powder. Recrystallization of the powder from acetic acid using active charcoal yielded III, long yellow needles, m. p. 299~304°C, 22 g. (48%). Further recrystallization of this material from acetic acid gave pure III, m. p. 304~306°C, (value in literature, m. p. 300.5°C[5]). III is soluble in concentrated sulfuric acid giving a greenish yellow solution.

Found: C, 83.69; H, 3.47; N, 12.13. Calcd. for $C_{16}H_8N_2$: C, 84.19; H, 3.53; N, 12.28%.

I. R. max., 2225 cm^{-1} (-C≡N).

1,8-Dicarboxyanthracene (VI).—*a) Hydrolysis of 1,8-Dicyanoanthracene (II).*—A mixture of II (10 g., 0.044 mol.), potassium hydroxide (125 g.) and ethylene glycol (270 ml.) was heated to 150~155°C under stirring until the evolution of ammonia ceased (ca. 70 hr.). The reaction mixture was neutralized with hydrochloric acid resulting in formation of a yellow precipitate. The precipitate was filtered, washed with water and dried, yielding yellow powder, 11 g. (94%), m. p. 341~345°C (decomp.). Recrystallization of this material from 95% ethanol resulted in pure VI, yellow needles, m. p. 347°C (decomp.). (value in literature, 345°C (decomp.)[5]).

Found: C, 71.84; H, 3.77. Calcd. for $C_{16}H_{10}O_4$: C, 72.18; H, 3.79%.

b) Reduction of 1,8-Dicarboxyanthraquinone (V).—1,8-Dicyanoanthraquinone (IV) which was obtained by the reaction of I with cuprous cyanide in benzyl cyanide[5b] was hydrolyzed to V by the reaction with sulfuric acid[5]. The reduction of V by means of zinc dust in aqueous ammonia yielded

** All melting points are uncorrected. The infrared spectra were measured by the Nujol mull method otherwise special indication is given.

10) It has been reported that the reduction of anthracene with zinc dust in aqueous ammonia yields anthranol. Cf., H. R. von Perger, *J. prakt. Chem.*, **23**, 137 (1881).

11) E. B. Barnett and M. A. Matthews, *J. Chem. Soc.*, **123**, 2549 (1923).

VI, m. p. 347°C (decomp.). This substance showed no depression of the melting point on admixture with the specimen prepared by route *a*).

Anthracene-1, 8-dicarboxylic Acid Chloride (VII).—The acid chloride (VII) was prepared according to the procedure of Waldmann[5b]. Recrystallization of the crude product from toluene gave pure VII, orange needles, m. p. 255~257°C (decomp., in a sealed capillary tube). (value in literature, m. p. lower than 220°C (decomp.)[5b]). VII gives yellow solution in concentrated sulfuric acid.

Found: C, 63.15; H, 2.80; Cl, 23.37. Calcd. for $C_{16}H_8O_2Cl_2$: C, 63.39; H, 2.66; Cl, 23.39%.

1, 8-Diacetylanthracene (IX).—1) *Malonation of the Acid Chloride (VII).*—Diethyl ethoxymagnesiomalonate was prepared from magnesium (11 g., 0.46 gram atom), diethylmalonate (73.5 g., 0.46 mol.) and absolute ethanol (27.6 g., 0.6 mol.) in dry xylene (100 ml.). The excess of ethanol was removed by the azeotropic distillation with xylene. The solution of the malonate was added to the hot solution of VII (17 g., 0.057 mol.) in dry xylene (1.8 l.) in 30 min. under stirring. The mixture was refluxed for 1 hr. and decomposed with the addition of crushed ice (100 g.) and 4 N sulfuric acid (300 ml.). The organic layer was washed with water, and dried over anhydrous sodium sulfate. Evaporation of the solvent under reduced pressure resulted in the crude ketomalonate (VIII) as an orange red oil.

2). *Ester Exchange and Decarboxylation of the Ketomalonate (VIII).*—The crude VIII was mixed with propionic acid (320 ml.) and concentrated sulfuric acid (4 ml.), and the mixture was refluxed for 4 hr. with a vigorous evolution of carbon dioxide. Sulfuric acid (4 N, 250 ml.) was added to the reaction mixture and the reflux was continued until no more evolution of carbon dioxide was observed. The mixture was distilled to remove a small amount of xylene with water, and poured onto ice-water (1 l.). The ice cooled mixture was filtered. The solid thus obtained was washed with water, and dissolved in benzene. The benzene solution was washed with dilute sodium hydroxide solution and water, successively. The dried solution was filtered through a short column of aluminum oxide, and the concentration of the filtrate resulted in IX, light yellow cubes, 7.4 g. (50%.), m. p. 176~179°C (Kofler block). Recrystallization of this substance from benzene resulted in pure IX, light yellow cubic crystals, m. p. 179~180°C. IX gives an orange red solution with concentrated sulfuric acid.

Found: C, 82.49; H, 5.49. Calcd. for $C_{18}H_{14}O_2$: C, 82.43; H, 5.38%.

I. R. max., 1675 cm^{-1} (C=O).

1, 8-Diethynylanthracene (IX).—*1) Chlorination of the Diacetylanthracene (IX).*—IX (5.25 g., 0.02 mol.) was mixed with phosphorus pentachloride (12.5 g., 0.06mol.) and phosphorus oxychloride (50 ml.), and the mixture was heated under reflux for 1.5 hr. With an evolution of hydrogen chloride, the mixture changed to a homogeneous dark red solution. The oxychloride was removed completely under reduced pressure. The dark red viscous oil was kept for several hours in a desiccator containing phosphorus pentoxide and potassium hydroxide.

2). *Dehydrochlorination of the Chloride (X).*—The crude chloride (X) in absolute ether (30 ml.) was added in 40 min. to a solution of sodium amide (prepared from sodium, 11.5 g., 0.5 gram atom) in liquid ammonia (500 ml.). The mixture was stirred for 2 hr. at the temperature of dry ice-ethanol bath. After evaporation of ammonia, water (250 ml.) was added cautiously to the reaction mixture and filtered to collect insoluble material. The dark brown solid obtained was digested with light petroleum (b. p. 60~80°C, 1.5 l.). The solution was washed with water and dried over anhydrous sodium sulfate. The solution was passed through a short column of alumina, and the filtrate was evaporated to yield XI, 1.65 g. (37%), m. p. 150~151°C (decomp., Kofler block). XI gave red cuprous, yellow silver and mercuric acetylides.

Found: C, 95.49; H, 4.45. Calcd. for $C_{18}H_{10}$: C, 95.54; H, 4.46%.

I. R. max., 3260 cm^{-1} (-C≡CH).

U. V. max. (in cyclohexane), λ_{max} 258 mμ (log ε =5.13), 279 (4.08), 291 (4.08), 312 (2.87)*, 330 (3.22), 345 (3.60), 362 (3.98), 381 (4.22) and 402 (4.24). Asterisk indicates shoulder.

Oxidative Coupling of 1, 8-Diethynylanthracene (XI).—To the solution of XI (0.600 g., 0.0027 mol.) in pyridine (100 ml.) and methanol (8 ml.) was added cupric acetate monohydrate (12 g.), and the mixture was vigorously stirred for 4 hr. at 55°C. Stirring was continued for a further 3 hr. period at room temperature. The reaction mixture was filtered and the solid obtained was washed thoroughly with water, a small amount of benzene and ethanol, successively. The cyclic tetrayne XII almost free from inorganic contaminate was obtained as orange powder, 0.562 g. (95%). This substance decomposed over 370°C. A minor amount of XII of poor purity was obtained from the pyridine solution.

The crude XII was recrystallized 3 times from xylene and once from dimethylformamide yielding tiny orange cubes. These crystals decomposed over 370°C in an evacuated capillary tube without any change of the crystal form.

Found: C, 95.72; H, 3.61. Calcd. for $C_{36}H_{16}$: C, 96.40; H, 3.60%.

The unsatisfactory result in the analysis of carbon may be due to explosive decomposition of XII in the combustion tube forming a carbonaceous material difficult to burn.

U. V. max., λ_{max} 258 mμ (log ε=5.22) (in *n*-hexane), 307 (4.82), 340 (3.66), 358 (3.79), 376 (3.93), 413 (4.22), 437 (4.67) and 465 (4.96) (in benzene).

Reduction of the Cyclic Tetraacetylene (XII) with Sodium in Liquid Ammonia.—The mixture of XII (0.42 g., 0.0009 mol.), sodium (0.9 g., 0.039 gram atom) and liquid ammonia (150 ml.) was stirred for 7 hr. at the temperature of dry ice-ethanol bath. Ammonium nitrate was added to the reaction mixture, and ammonia was evaporated. Water was added to the residue and the light yellow powder formed was collected by filtration. The powder was digested with benzene to remove the insoluble material. The benzene solution was filtered

through a thin layer of alumina (10 g.). Evaporation of the solvent afforded XIII, yellow needles, 0.20 g. (48%), m. p. 365~368°C (decomp., in an evacuated capillary tube, metal block).

Found: C, 93.24; H, 6.67. Calcd. for $C_{36}H_{30}$: C, 93.46; H, 6.54%.

I. R. max. (KBr-disk method), 2920, 2860 ($-CH_2-$), 952 (isolated trans-ethylenic bond), 875 cm^{-1} (tetra-substituted benzene).

U. V. max., λ_{max} 257 mμ ($\log \varepsilon = 4.91$) (in *n*-hexane), 318 (3.71), 332 (3.72), 348 (3.73), 368 (3.84) and 387 (3.89) (in benzene).

Catalytic Reduction of $C_{36}H_{30}$ (XIII).—The solution of XII (0.19 g.) in benzene (300 ml.) was shaken in hydrogen with palladium (5%) on charcoal (0.5 g.). The color and the fluorescence of the solution gradually disappeared in the course of the hydrogenation. The catalyst was filtered and washed with benzene. The filtrate was combined with the washing, and the evaporation of the solvent yielded XIV as fine colorless needles in almost quantitative yield. This was recrystallized from benzene to give pure XIV, m. p. 316~318°C, (in an evacuated capillary tube; metal block).

Found: C, 92.10; H, 7.67; Mol. wt. (Rast, camphor), 450. Calcd. for $C_{36}H_{36}$: C, 92.26; H, 7.74%; Mol. wt., 469.

I. R. max. (KBr-disk method), 2935, 2870 ($-CH_2-$), 735 cm^{-1} ($-(CH_2)_4-$).

U. V. max. (in ethanol), λ_{max} 256 mμ; λ_{min} 248 mμ.

The financial support of this research was granted by the Ministry of Education. The starting material was donated by the Mitsui Chemical Industries Co., Ltd. The elementary analyses were performed by Mr. M. Okumiya and Miss K. Koike. The authors wish to express their sincere thanks for these donation and for the help received.

Department of Chemistry
Faculty of Science
Osaka University
Nakanoshima, Osaka

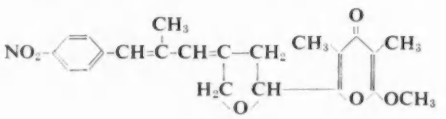
The Structure of Aureothin. X[1]*. Aza Derivatives of Aureothin, Isoaureothin and Desmethylisoaureothin*

By Kiyoyuki YAMADA, Hisao NAKATA and Yoshimasa HIRATA

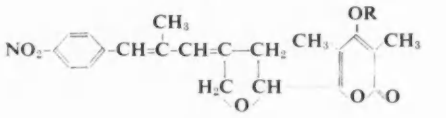
(Received April 2, 1960)

A recent publication[2] from our laboratory has indicated that the structure of aureothin was shown to be I, on a firm basis of a quantity of chemical and spectroscopic evidence. Aureothin undergoes demethylation with various acids, followed by isomerization, to give desmethylisoaureothin (II, R=H), the methylation of which, in turn leads to isoaureothin (II, $R=CH_3$) as a major product[1].

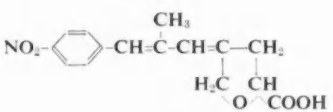
During the course of the structural investigation, it was found that I, II ($R=CH_3$) and II (R=H) reacted with ammonia and were transformed to aza-aureothin (IV), aza-isoaureothin (V) and aza-desmethylisoaureothin (VI), respectively. In the present paper, the properties of these aza-derivatives are described, being compared with those of other model compounds. Each of these aza-derivatives contains one more nitrogen, besides an original one (nitro group), and is sparingly soluble both in organic solvents and water.



(I)



(II)



(III)

Aza-aureothin (IV).—The elemental analysis of IV was incompatible with the molecular formula, expected for γ-pyrone—γ-pyridone

1) For Part IX of this series, see, H. Nakata, This Bulletin, in press.

2) K. Yamada, H. Nakata and Y. Hirata, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **81**, 340 (1960).

3) Y. Hirata, K. Okuhara, H. Nakata, T. Naito and K. Iwadare, ibid., **78**, 1700 (1957).

transformation; there remained no methoxyl group in IV. Thus, somewhat complicated reactions must take place during the formation of IV from I. It is apparent from the ultraviolet spectrum that γ-pyrone moiety is involved in the reaction, and the conjugated *p*-nitrophenyl system remains intact. Subtraction of the ultraviolet spectrum of aureothinic acid (III) from that of aza-aureothin (IV) gave an absorption spectrum due to the chromophore in question—presumably, a nitrogen-containing ring; the subtraction curves, taken in neutral and acid solutions, were different, as shown in Fig. 1.

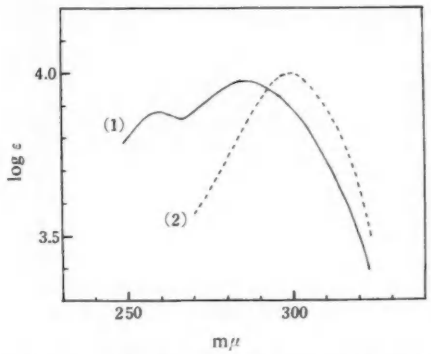


Fig. 1. Substraction curves obtained by aza-aureothin (IV)—aureothinic acid (III) in ethanol.
(1): Neutral and alkaline (0.1 N EtONa) solutions.
(2): Acidic (0.1 N HCl) solution.

The infrared spectrum showed bands at 1666(s), 1643(m), 1583(s), 1529(s), 1509(s) cm^{-1} and at 3385, 3285 and 3175 cm^{-1}, the latter group of bands suggesting the presence of N–H and/or O–H. Treatment of IV with acetic anhydride and pyridine yielded an amorphous substance, the infrared spectrum of which showed bands at 1731(s), 1666(s), 1623(s), 1596(s) and 1516(s) cm^{-1} (in chloroform). This must be an ester. The insolubility of IV in various solvents prevented us from attempting further studies on this substance, and the details of the structure are still uncertain.

Aza-isoaureothin (V).—Analytical data of V agreed with a molecular formula $C_{21}H_{22}O_5N_2$, identical with that of aza-desmethylisoaureothin (VI), but the chemical behavior, as described below, is quite different from those expected for aza-desmethylisoaureothin (VI). In this case, again, some other reaction might occur concomitantly with the pyrone—pyridone transformation. This compound is not an acidic substance but a neutral one, and an attempted acetylation only resulted in the complete recovery of the starting material. The same procedure used for IV was applied for the measurement of the ultraviolet spectrum; the results obtained by the subtraction method of the ultraviolet curves are shown in Fig. 2. In this case also, the absorption maximum in an acidic medium is situated at somewhat longer wavelength than that in neutral solution.

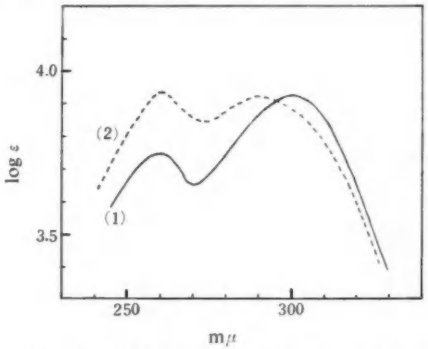


Fig. 2. Subtraction curves obtained by aza-isoaureothin (V)—aureothinic acid (III) in ethanol.
(1): Neutral and alkaline (0.1 N EtONa) solutions.
(2): Acidic (0.1 N HCl) solution.

Remarkable differences were observed in the infrared spectra between isoaureothin (II, $R=CH_3$) and aza-isoaureothin (V), as illustrated in Fig. 3. The following points are worth noticing; the shift of carbonyl band (1716 cm^{-1} for isoaureothin, 1660 cm^{-1} for aza-isoaureothin), and the appearance of a new band at 1554 cm^{-1} and of a series of bands at 3225, 3322, 3434 cm^{-1}.

From these spectroscopic data, it is seen that an α-pyridone ring is likely to be present in V, but the whole structural aspect of V could not yet be disclosed, owing to the same reason, as that encountered in the case of aza-aureothin (IV).

Aza-desmethylisoaureothin (VI).—Aza-desmethylisoaureothin $C_{21}H_{22}O_5N_2$, was prepared via two routes; one is based on a reaction of desmethylisoaureothin (II, R = H) with ammonia, and the other is due to the treatment of aureothin with ammonium acetate in acetic acid. Chemical and spectroscopic behavior of VI, described below, substantiates a simple pyrone-pyridone transformation. VI dissolved in aqueous alkali, although no reaction took place between VI and ethereal diazomethane (under heterogeneous condition). Acetylation of VI with acetic anhydride-pyridine gave aza-desmethylisoaureothin acetate, $C_{23}H_{24}O_6N_2$ (VII), which showed an enol acetate band at 1767 cm^{-1} (KBr).

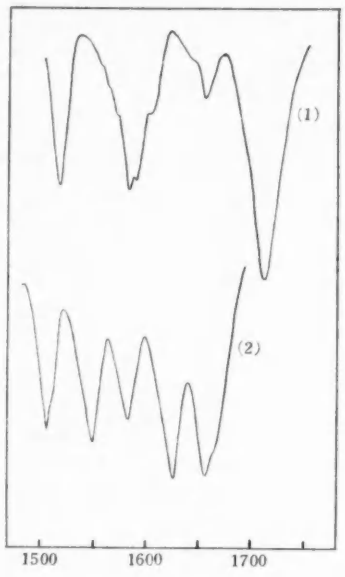


Fig. 3. Infrared spectra (KBr)
(1): Isoaureothin. (II, R=CH_3).
(2): Aza-isoaureothin. (V).

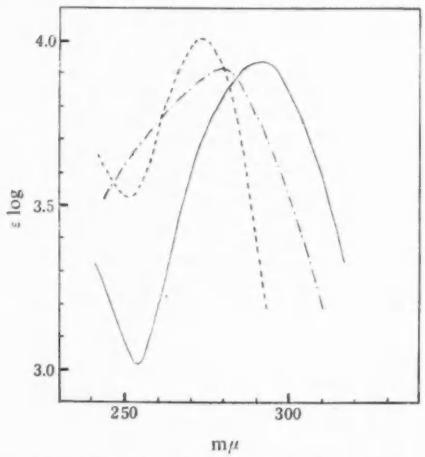
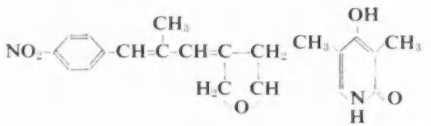


Fig. 4. Subtraction curves obtained by aza-desmethylisoaureothin (VI)—aureothinic acid (III) in ethanol.
(1): Neutral solution.
(2): Acidic (0.1 N HCl) solution.
(3): Alkaline (0.1 N EtONa) solution.

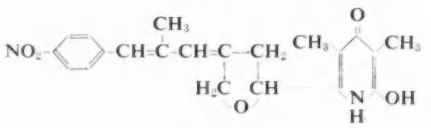
In the infrared spectrum of VI, there appeared bands at 1643(s), 1605(v. s), 1585(v. s), 1564(m) and 1510(s) cm^{-1} in 1800~1500 cm^{-1} region and a broad band ranging from 2500 cm^{-1} to 3000 cm^{-1}, which indicates the occurrence of hydrogen bonding. The position of pyridone-carbonyl band unaffected by hydrogen bonding is available from aza-desmethylisoaureothin acetate (VII); i.e., in this case, $\nu_{C=O}$ appeared at 1654 cm^{-1} (KBr).

In the ultraviolet region, the absorption due to the pyridone ring of VI had its maximum at 291 mμ in neutral solution, while acidic and alkaline conditions caused the hypsochromic displacement of the absorption (Fig. 4). The moderate solubility in chloroform made possible the ozonization of VII, which afforded aza-desmethylisoaureonone acetate, $C_{15}H_{15}O_5N$ (VIII). This product VIII, now free from the conjugated *p*-nitrophenyl system, showed an absorption maximum at 304 mμ in the ultraviolet region, and bands at 1762(s), 1658(s), 1640(sh.) and 1577(w) cm^{-1} in the infrared region.

Structure of Aza-desmethylisoaureothin. — Both types VIa and VIb would be possible for VI, since enolization can occur in two directions[4].



(VIa)



(VIb)

A considerable amount of data[5] concerning the infrared spectra of various α-pyridones is now at hand. The condition of the measurement (solid or solution) has a fairly large influence on the position of carbonyl absorption. The ranges quoted for carbonyl absorption in the solid state are rather wide (1630~1680 cm^{-1})[5], whereas in solution, it has been reported that carbonyl frequencies fell into a somewhat narrow range (1654~1687 cm^{-1})[6].

As to γ- pyridones, the data so far are not sufficient to obtain the detailed information about $\nu_{C=O}$; in general, $\nu_{C=O}$ is expected to appear in

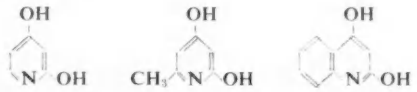
4) One more possibilty, the diketo-form, is excluded on the ground of its relation to the ultraviolet spectrum.

5) (a) W. F. Cockburn and L. Marion, *Can. J. Chem.*, **29**, 13 (1951). (b) F. Ramirez and A. P. Paul, *J. Org. Chem.*, **19**, 183 (1954); *J. Am. Chem. Soc.*, **77**, 1035, 3377 (1955). (c) J. A. Gibson, W. Kynaston and A. S. Lindsey, *J. Chem. Soc.*, **1955**, 4340. (d) F. Bohlman, A. Englisch, N. Ottawa, H. Sander and W. Weise, *Chem. Ber.*, **89**, 792 (1956). (e) K. Hoegerle and H. Erlenmeyer, *Helv. Chim. Acta*, **39**, 1203 (1956). (f) R. H. Wiley and S. C. Slaymaker, *J. Am. Chem. Soc.*, **78**, 2393 (1956).

6) S. F. Mason, *J. Chem. Soc.*, **1957**, 4874.

the region 1630~1650 cm^{-1} (solid and solution)[7]. According to Mason, the C=O band of the α-pyridones is located at higher frequency than that of γ-isomers, whilst the reverse relation is observed regarding the N-H band[6]. Further, it was noticed[7d] that in α-pyridones, the intensity of $\nu_{C=O}$ is usually stronger than that of $\nu_{C=C}$, whereas in γ-pyridones carbonyl absorption appears as a relatively weak band.

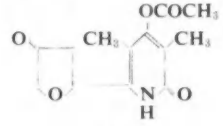
The following compounds may serve as good models of aza-desmethylisoaureothin (VI), in relation to the problem of tautomerism; 2, 4-dihydroxypyridine (IX)[6], 6-methyl-2, 4-dihydroxypyridine (X) and 2, 4-dihydroxyquinoline (XI)[6].



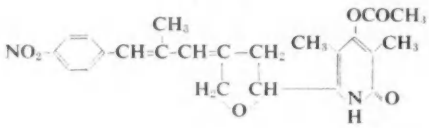
(IX) (X) (XI)

As to dihydroxypyridine (IX) it was believed that an α-pyridone form was predominant in aqueous solution[8] on the ground of the ultraviolet evidence, while the possible coexistence of α, γ-forms was suggested for IX in the solid state[9]. As for 6-methyl-2, 4-dihydroxypyridine (X), it is insufficiently soluble in chloroform to give an infrared spectrum in solution, but in the solid state (KBr), there appeared bands at 1658(s), 1639(s) and 1607(s) cm^{-1} and a broad band near 3000 cm^{-1} indicating a hydrogen bond. With the information about the diagnosis of α, γ-tautomers at hand, it is concluded that the tautomer of α-pyridone form occurs in X. An α-pyridone structure was inferred[7d] for XI itself and its acetate, but not established firmly as yet.

With the help of the generalization described above, the structural examination of aza-desmethylisoaureothin (VI) will be made. The structure VIII is assigned to aza-desmethylisoaureonone acetate, the ozonisation product, on the basis of the position (1658 cm^{-1}) and the strong intensity of the pyridone carbonyl band. A strong band at 1762 cm^{-1} in VIII is due to two carbonyls of enol acetate and five-membered ketone of tetrahydrofuran ring; two bands at 1640 and 1557 cm^{-1} are ascribed to the unsaturation of the pyridone ring[10]. Establishment of VIII for the ozonisation product leads to structure VII for aza-desmethylisoaureothin acetate.



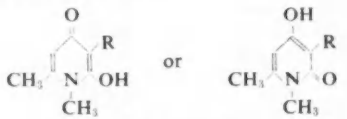
(VIII)



(VII)

The disagreement in the position of the carbonyl band of aza-desmethylisoaureothin (1643 cm^{-1}) with that of the model compound (X, 1658 cm^{-1}), taken in conjunction with the weak intensity in the former $\nu_{C=O}$, might be in favor of the γ-pyridone structure VIb for this derivative of the natural product. However, in aza-desmethylisoaureothin, the intermolecular (or intramolecular) interaction would be more complicated than in a simple molecule, and the overlappings of the bands due to benzene ring with those of the pyridone ring make it difficult to compare exactly the intensities between $\nu_{C=O}$ and $\nu_{C=C}$ in the pyridone ring. Such circumstances give rise to difficulty in making the discrimination between (Va) and Vb, solely on the basis of the infrared spectrum.

The unusual hypsochromic shift of the ultraviolet absorption maximum in alkaline solution was recognized as a characteristic feature of the hydroxypyridones (A), distinguishing the type (A) from other types of hydroxypyridone[11]. As is expected, the model compound IX behaves similarly (see Fig. 4).



(A): R=H or C_2H_5

A good paralellism of the ultraviolet spectra between the pyridone part of VI and the model X was observed, in regard to the magnitudes

7) (a) R. Adams, V. V. Jones and J. L. Johnson, *J. Am. Chem. Soc.*, **69** 1810 (1947); (b) N. J. Leonard and D. M. Locke, ibid., **77**, 1852 (1955); (c) P. I. Ittyeran and F. G. Mann, *J. Chem. Soc.*, **1958**, 467; (d) H. Shindo, "Infrared Absorption Spectrum Series No. 3" Nankodo, Tokyo, (1957), p. 186.

8) H. J. den Hertog and D. J. Buurman, *Rec. trav. chim.*, **75**, 257 (1956). Cf. H. M. Woodburn and M. Hellman, ibid., **70**, 813 (1951).

9) There was recorded in the solid state, a pair of bands, 1661 and 1635 cm^{-1} for IX (1662 and 1628 cm^{-1} for (XI)), upon which the presumption of the coexistence of α, γ-tautomers was based. Alternative interpretation of the spectrum, however, seems to be more reasonable; both IX and XI exist simply as an α-pyridone form on the assumption that bands, 1635 and 1628 cm^{-1} are respectively, associated with the unsaturated bonds of the ring.

10) The positions of bands associated with the unsaturated bonds of the ring are very sensitive to the substituents[5f]; therefore, these bands are not so useful for the identification of the pyridone ring.

11) J. A. Berson, W. M. Jones and S. L. F. O'callaghan, C. S. J., *J. Am. Chem. Soc.*, **78**, 622 (1956).

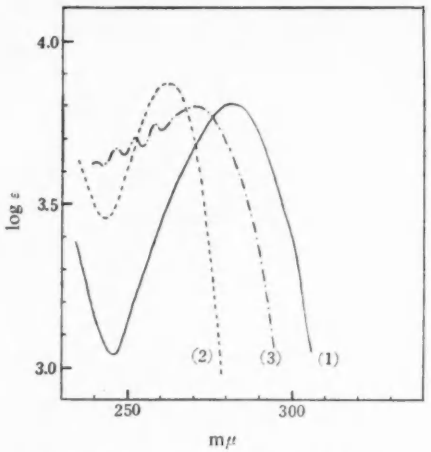


Fig. 5. Ultraviolet spectra of 6-methyl-4-hydroxy-2-pyridone (X) in ethanol.
(1): Neutral solution.
(2): Acidic (0.1 N HCl) solution.
(3): Alkaline (0.1 N EtONa) solution.

of the hypsochromic shifts and of the relative intensities. It is inferred that α-pyridone form VIa predominates in aza-desmethylisoaureothin in solution, on the analogy of the preferential α-tautomer in the cases of the models IX and X. Furthermore, the absorption at 292 mμ due to the pyridone part of VI is regarded as a reasonable position of α-pyridone ring[12].

Experimental

Aza-aureothin (IV).—One gram of aureothin was added in a mixture of ethanol (40 ml.) and aqueous ammonia (10 ml.), and the suspension was kept at 90~95°C in a sealed tube for 10~12 hr. Aureothin gradually dissolved with heating, and after several hours a small number of crystals began to appear. The solution on cooling to room temperature deposited yellow crystals, which were filtered and recrystallized from a large amount of ethanol, m. p. 230~231°C. No methoxyl group was detected by Zeisel's method; the ultraviolet spectrum was measured in neutral ethanol, by assuming 380 for its molecular weight: λ_{max} 345 mμ (log ε 4.25), 290 mμ (log ε 4.15).

Found: C, 64.75; H, 5.68%.

A suspension of 20 mg. of aza-aureothin was heated in acetic anhydride (5 ml.) and pyridine (1 ml.) at 80~90°C for thirty minutes. Water was added, to give the yellow precipitate, easily soluble in chloroform and ethanol.

Aza-isoaureothin (V).—A suspension of 1 g. of isoaureothin in 50 ml. of ethanol-aqueous ammonia (4 : 1) was heated at 90~95°C in a sealed tube for 18 hr. Complete dissolution of isoaureothin was observed within the initial five hours, and then crystallization of V slowly began to take place. As the heating was continued, the solution became dark red. The product was collected and treated with 20 ml. of hot ethanol, in order to remove remaining isoaureothin. Recrystallization from a fairly large amount of dioxane gave 360 mg. of V in yellow prisms, m. p. 228~229°C. The ultraviolet spectrum was measured in neutral ethanol with the assumption of 380 of molecular weight: λ_{max} 343 mμ (log ε 4.25), 253 mμ (log ε 4.25).

Found: C, 65.84, 65.62; H, 5.99, 5.97; N, 7.33%.

A mixture, which consisted of 20 mg. of V, 8 ml. of acetic anhydride and 1 ml. of pyridine, was heated at 80~90°C for 1 hr. to give a clear solution. Upon cooling, the yellow needles separated. These were identified as starting material V.

Aza-desmethylisoaureothin (VI).—Though aza-desmethylisoaureothin was obtained by a procedure similar to that used for the preparation of IV and V, an efficient method, starting from aureothin, was developed, to give an excellent yield of (VI). A suspension of 1 g. of aureothin in 10 ml. of acetic acid containing 3 g. of ammonium acetate was gently refluxed for 4 hr. To the cooled solution, 200 ml. of water was added, when a brown oily layer at once separated and solidified. The liquid layer was removed by decantation, and the residue turned to a crystalline substance by the addition of 50ml. of hot ethanol. Washed with hot ethanol three times, the product was recrystallized from dimethylformamide and weighed 250 mg., m. p. 238~239°C; ultraviolet spectrum: λ_{max} 341 mμ (log ε 4.25), 296 mμ (log ε 4.16), 245 mμ (log ε 4.18) (ethanol as solvent).

Found: C, 65.76, 65.48; H, 5.95, 6.19; N, 7.63. Calcd. for $C_{21}H_{22}O_5N_2$: C, 65.95; H, 5.80; N, 7.33%.

To a solution of 1 ml. of pyridine and 5 ml. of acetic anhydride, 20 mg. of VI was suspended. The mixture became a clear solution when kept at 70~80°C for 30 min. Yellow precipitates were obtained from the cooled solution, by the addition of 20 ml. of water. They were recrystallized from ethanol whereupon yellow needles of VII with m. p. 211~213°C were obtained (15 mg.): ultraviolet spectrum: λ_{max} 325 mμ (log ε 4.32), 238 mμ (log ε 4.30); infrared spectrum: 1767(v. s), 1652(s), 1610(m), 1591(s) and 1516(s)cm^{-1} (KBr).

Aza-desmethylisoaureonone Acetate (VIII).—A solution of 350 mg. (0.83 m. mol.) of aza-desmethylisoaureothin acetate (VII) in 15 ml. of chloroform was cooled to −5~−10°C and ozone was passed through this solution, until about one equivalent amount (0.95 m. mol.) of ozone was absorbed. On removal of solvent under reduced pressure, a gummy residue was obtained, which was treated with 10 ml. of water in a flask provided by a reflux condenser at 60°C. The aqueous layer in the flask was separated from the solid residue by decantation and the aqueous fraction was distilled under reduced pressure to afford a slightly yellow residue, which on recrystallization from methanol, weighed 10 mg.

12) The ultraviolet technique is not necessarily reliable for the discrimination of α-pyridones from their γ-isomers, in spite of the fact that α-pyridone itself absorbs at a much higher wave length (297 mμ) than γ-isomer (256 mμ); the substitution of the ring often causes an outstanding change in the spectrum.

with m. p. 208~209°C. No volatile carbonyl compound could be detected by distilling the aqueous fraction into a 2, 4-dinitrophenylhydrazine solution. The solid residue, obtained from the decomposition of ozonide, consisted of two components; one was easily soluble in methanol, whereas the other was sparingly soluble. The former was identified as *p*-nitro-α-methylcinnamaldehyde by its melting point and infrared spectrum. The latter, sparingly soluble in methanol, was recrystallized from methanol and gave 70 mg. of colorless needles, which turned out to be identical with the crystals obtained from the water soluble fraction, as described before. The total amount was 80 mg., m. p. 209~210°C; ultraviolet spectrum: λ_{max} 304 mμ (log ε 3.85), 233 mμ (log ε 3.81).

Found: C, 58.76, 58.86; H, 5.84, 5.98; N, 5.66, 5.45. Calcd. for $C_{18}H_{15}O_5N$: C, 58.83; H, 5.70; N, 5.28%.

Summary

Aureothin (I), isoaureothin (II, R=CH_3) and desmethylisoaureothin (II, R=H) reacted with ammonia in a somewhat drastic way, and gave aza-derivatives respectively, some properties of which are described. Of these, aza-desmethylisoaureothin was shown to be VI containing tautomeric α, γ-dihydroxypridine, on a basis of chemical and spectroscopic evidence.

The authors are indebted to Dr. Koichi Iwadare, Banyu Pharmaceutical Co. for his generous supply of aureothin used in this research, and to Mr. Takayuki Naito, Banyu Pharmaceutical Co., and Mr. Kunio Okuhara, Government Industrial Research Institute of Nagoya, for their helpful discussion throughout this work. The authors also wish to thank the Ministry of Education for a grant-in-aid.

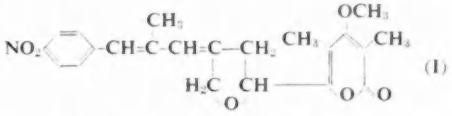
Chemical Institute
Faculty of Science
Nagoya University
Chikusa-ku, Nagoya

The Structure of Aureothin. XI[1]. *On the Structure of Some of the Isoaureothin Derivatives*

By Kiyoyuki YAMADA, Takayuki NAITO, Kunio OKUHARA, Hisao NAKATA and Yoshimasa HIRATA

(Received April 2, 1960)

Among many derivatives obtained from isoaureothin (I)[2], there are several, the structures of which has not yet been disclosed. These derivatives, owing to their unexpected behavior, made the investigation of isoaureothin rather complicated. In this paper, the structures of these derivatives are put forward. They are well related to isoaureothin (I) and, of course account for all their properties. Further, the aspect of these reactions is briefly discussed. Of these derivatives, some descriptions were already made as to the nature of the products obtained by complex metal hydride reduction[3].



(I)

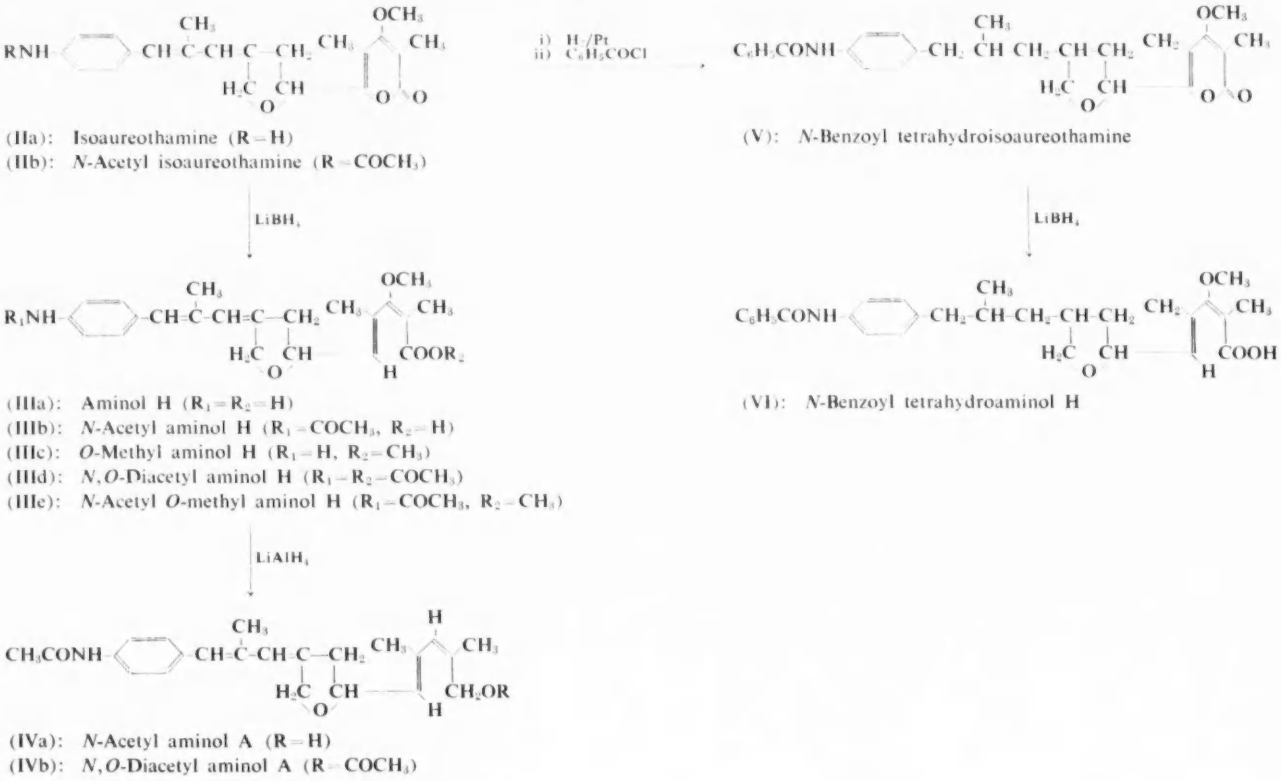
1) Part X. K. Yamada, H. Nakata and Y. Hirata, This Bulletin, **33**, 1298 (1960).
2) K. Yamada, H. Nakata and Y. Hirata, *J. Chem. Soc. Japan, Pure Chem. Sec.* (*Nippon Kagaku Zasshi*), **81**, 340 (1960).
3) K. Yamada, Y. Hirata, K. Okuhara, H. Nakata, T. Naito and K. Iwadare, ibid., **79**, 384 (1958).

On the Structures of Aminol Compounds

Structure of Aminol H (IIIa).—The reduction of the nitro group in isoaureothin (I) by ferrous hydroxide afforded a corresponding amino derivative $C_{22}H_{25}O_4N$ (IIa), which in turn was treated with lithium borohydride in tetrahydrofuran to give an acidic compound named "aminol H", $C_{22}H_{27}O_4N$[3]. As regards the acidic nature of aminol H, the infrared evidence (broad band at 2600~2700 cm^{-1} in aminol H: bands at 1795, 1712 cm^{-1} in *N, O*-diacetyl aminol H (IIId)), coupled with the result of the methoxyl analysis[4], showed that it might be due to a carboxyl group. *p*-Bromophenacyl ester of aminol H (m. p. 85~87°C) was obtained. Taking account of the fact that aminol H is optically active and has two more hydrogen atoms than isoaureothamine (IIa), it seems reasonable to deduce the occurrence of the hydrogenolytic cleavage in α-pyrone ring,

4) Aminol H (IIIa) contains one methoxyl group, whereas two such groups present in the methylated product, *O*-methyl aminol H (IIIc).

Chart 1. Derivatives of isoaureothin.



(IIa): Isoaureothamine (R=H)
(IIb): *N*-Acetyl isoaureothamine (R=$COCH_3$)

(V): *N*-Benzoyl tetrahydroisoaureothamine

(IIIa): Aminol H ($R_1=R_2=H$)
(IIIb): *N*-Acetyl aminol H ($R_1=COCH_3$, $R_2=H$)
(IIIc): *O*-Methyl aminol H ($R_1=H$, $R_2=CH_3$)
(IIId): *N,O*-Diacetyl aminol H ($R_1=R_2=COCH_3$)
(IIIe): *N*-Acetyl *O*-methyl aminol H ($R_1=COCH_3$, $R_2=CH_3$)

(VI): *N*-Benzoyl tetrahydroaminol H

(IVa): *N*-Acetyl aminol A (R=H)
(IVb): *N,O*-Diacetyl aminol A (R=$COCH_3$)

Table I. Yield of methyl pyruvate

Compound	Sample mg.	$CH_3COCOOCH_3$[a] mg.	Equiv. mol.
Isoaureothin (I)	940	130	0.19
N-Acetyl isoaureothamine (IIb)	300	40	0.21
N-Benzoyl tetrahydroisoaureothamine (V)	470	32	0.15
N-Acetyl aminol H (IIIb)	250	38	0.22
N-Benzoyl tetrahydroaminol H (VI)[b]	620	80	0.22
N-Acetyl *O*-methyl aminol H (IIIe)	450	311	1.05

a) Determined as 2,4-dinitrophenylhydrazone.
b) Crude sample (m. p. 78~83°C) was used without recrystallisation.

to produce a carboxyl group[5]. The various properties previously described[3] (infrared and ultraviolet spectra etc.), thus can best be accounted for by assigning the structure IIIa to aminol H. Furthermore, the following experiment supports the correctness of formula IIIa; ozonisation of compounds belonging to the isoaureothin- and aminol-series always afforded methyl pyruvate, the yield of which in each case is listed in Table I. An essentially equal amount of methyl pyruvate (ca., 0.2 equivalent mole) was produced in every case[6], except the largely enhanced amount (1.05 equivalent moles) in the case of *N*-acetyl *O*-methyl aminol H (IIIe). Therefore, in *N*-acetyl *O*-methyl aminol H (IIIe), it becomes evident that there exist in one molecule two such units, from which methyl pyruvate is formed: in addition, one of such units should originate from the newly formed carbomethoxyl group by complex metal hydride reduction and subsequent methylation (cf. methoxyl analysis[4]).

Structure of N-Acetyl Aminol A (IVa).—Reduction of *N*-acetyl *O*-methyl aminol H (IIIe) with lithium aluminum hydride gave a neutral product, *N*-acetyl aminol A, $C_{23}H_{25}O_2N$ in poor yield[3]; no other crystalline products than this were isolated. Elemental and methoxyl analysis showed, during the reduction of IIIe, the occurrence of unusual hydrogenolysis of methoxyl group (enol ether), originally present in isoaureothin (I), in addition to the expected transformation of the ester group into alcohol.

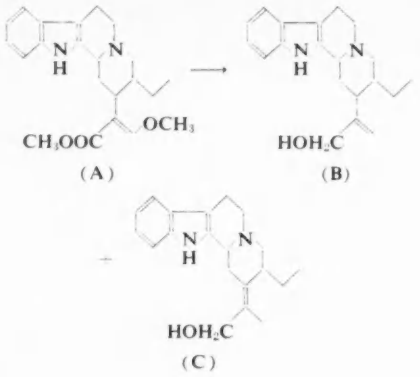
Aspect of Complex Metal Hydride Reduction.—As described above, plausible structural assignments were made to the reduction products, aminol H (IIIa) and *N*-acetyl aminol A (IVa), on a basis of their chemical and spectroscopic properties. These would, at first sight, be regarded as unusual reduction products.

In the case of aminol H, however, it is expected that α-pyrone shows peculiar reactivity far different from that of lactone toward hydride reagents, owing to the high conjugation of α-pyrone and a possible effect of substituents of the ring. So far, very little is known about the reduction of α- and γ-pyrones with various complex metal hydrides, except for limited examples of xanthone[7], benzo-α-[8] and benzo-γ-pyrones[9], where not unexpected results were recorded. In contrast to the reduction of such benzopyrones, a situation, however, would be unique in the case of pyrones if no aromatic ring was fused to them. Investigation on the reactivity of pyrones toward complex metal hydrides is desirable.

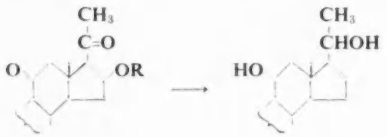
Of particular interest is a hydrogenolytic cleavage of enol-ether (methoxyl group), during the formation of *N*-acetyl aminol A (IVa). In general, enol ethers are inert to complex metal hydride reduction[10], and so is a β-alkoxy-α, β-unsaturated ester[11]. An exception[12] was reported; reduction of dihydrocorynantheine (A) with lithium aluminum hydride led to demethoxylated products B and C. Sarett and his coworkers reported a similar exception[13], found in β-alkoxy-α, β-unsaturated

5) On catalytic hydrogenation, aminol H absorbed quantitatively two moles of hydrogen, with which the double bonds conjugated with aminophenyl group were saturated[2]. From this result, it is tempting to consider the possibility of the existence of an α-pyrone ring in aminol H. However, this view is untenable on the ground that a band near 1575 cm^{-1}, characteristic of α-pyrone, is absent in all of the aminol compounds studied.
6) Ozonisation was carried out repeatedly, in order to ascertain the reproducible yield of methyl pyruvate in each case.

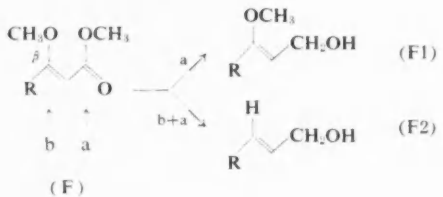
7) A. Mustafa and M. K. Hilmy, *J. Chem. Soc.*, **1952**, 1343.
8) (a) F. A. Hochstein, *J. Am. Chem. Soc.*, **71**, 305 (1949); (b) P. Karrer and P. Banerjea, *Helv. Chim. Acta*, **32**, 1692 (1949). (c) S. Siegel and S. Coburn, *J. Am. Chem. Soc.*, **73**, 5494 (1951); (d) J. P. Freeman and M. F. Hawthorne, ibid., **78**, 3366 (1956).
9) (a) R. Robinson and R. Mirza, *Nature*, **166**, 929 (1950); (b) R. Mirza and R. Robinson, ibid., **166**, 997 (1950); (c) M. Miyano and M. Matsui, *Chem. Ber.*, **91**, 2044 (1958).
10) N. G. Gaylord, "Reduction with Complex Metal Hydrides", Interscience Publishers, Inc., New York (1956), p. 695.
11) (a) J. C. Seaton and LeoMarion, *Can. J. Chem.*, **35**, 1102 (1957); (b) J. C. Seaton, R. Tondeur and LeoMarion, ibid., **36**, 1031 (1958).
12) P. Karrer, R. Schwyzer and A. Flam, *Helv. Chim. Acta*, **35**, 851 (1952); cf. A. Chatterjee and P. Karrer, ibid., **33**, 802 (1950).
13) G. E. Arth, G. I. Poos and L. H. Sarett, *J. Am. Chem. Soc.*, **77**, 3834 (1955).



ketone (D), which afforded product E lacking in alkoxyl group. Such an apparent anomaly



of β-alkoxyl group in α, β-unsaturated carbonyl system, encountered in the exceptions, can be interpreted in terms of vinylogy in reaction (see F); since β-carbon is regarded as a vinylogous position of carbonyl carbon, the reactivity of two methoxyl groups would be expected to be comparable.



Therefore, the mode of attack of the hydride reagent would be possible in two directions, as indicated by arrows in F, to give a normal product F1 and/or anomalous product F2. The subtle difference of the reactivity between two such positions would be influenced by structural variations and experimental conditions, neither of which, however, have yet been examined systematically.

Structure of "Lactol"

Property of Lactol (VIII). — *N*-Acetyl isoaureothamine (IIb) was converted to Lactol $C_{24}H_{29}O_6N$ (VIII), by the action of aqueous alkali[14]. The product, Lactol, is soluble in alkali and optically inactive, while the starting material IIb shows remarkable optical activity, $[\alpha]_D^{28^\circ}$ +140°.

The ultraviolet spectrum of Lactol shows the absence of α-pyrone chromophore (Fig. 1); the subtraction of ultraviolet spectrum of conjugated *p*-acetylaminophenyl group from that of Lactol gives a curve, which shows its maximum near 248 mμ and should correspond to that of α-pyrone ring (λ_{max} 295 mμ) of *N*-acetyl isoaureothamine (IIb). This outstanding hypsochromic shift, observed in the transformation of IIb to Lactol proves the destruction of α-pyrone chromophore in the latter[15]. There were observed bands at 1708(sh.), 1675 (s), 1645(s), 1599(s), 1530(s) and 1516(s)cm^{-1} in the 1800~1500cm^{-1} range of the infrared spectrum (KBr). The presence of a hydroxyl

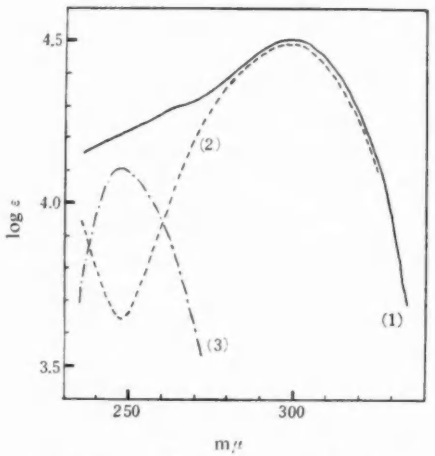
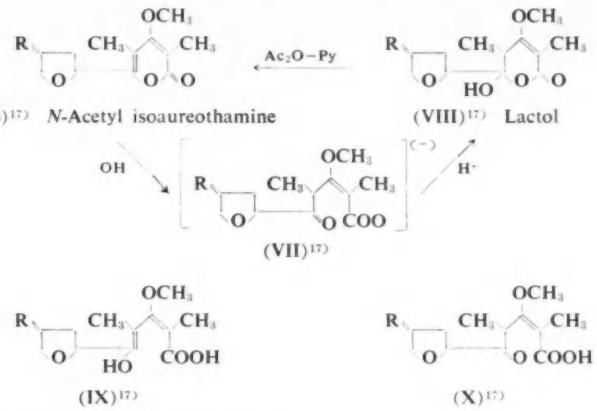


Fig. 1. Ultraviolet spectra (in ethanol)
(1): Lactol
(2): $CH_3CONH-C_6H_4-CH=C(CH_3)-CH=CH-CH=CH-OH$
(3): Subtraction curve, (1) — (2)

group in Lactol was detected from a broad band near 3200 cm^{-1} and a band at 1150 cm^{-1}. Lactol was neutralized by sodium hydroxide[16], to afford its sodium salt with bands located at 1717, 1672, 1640, 1595(s. and br.), 1530 and 1510 cm^{-1} (KBr). In order to clarify the nature of this hydroxyl group, acetylation with pyridine-acetic anhydride was carried out; an unexpected product, racemic *N*-acetyl isoaureothamine (IIb) was obtained. Methylation

14) In the same treatment of isoaureothin (I) and isoaureothamine (IIa), amorphous products were obtained showing properties similar to that of Lactol.
15) This view was also supported by the lack of the characteristic band of α-pyrone at 1575 cm^{-1} in Lactol.
16) The addition of excessive sodium hydroxide did not significantly after the infrared spectrum of the salt.



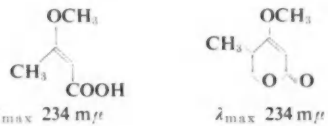
(IIb)[17] *N*-Acetyl isoaureothamine (VIII)[17] Lactol

(VII)[17]

(IX)[17] (X)[17]

with diazomethane brought out the same result obtained in acetylation. In spite of careful scrutiny of the reaction products, no methylated substance was obtained in crystalline state. Since the formation of sodium salt suggests an acidic nature of Lactol, measurement of acidity constant (pK_a) by titration method was undertaken. No consumption of alkali by sample was observed below pH 10.5, implying the absence or masked state of carboxyl group under the conditions studied.

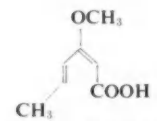
Structure of Lactol.—By comparison of the molecular formula between Lactol $C_{21}H_{29}O_6N$ and *N*-acetyl isoaureothamine $C_{21}H_{27}O_5N$, it is at once evident that hydrolysis of α-pyrone took place, leading to a product soluble in alkali. Since alcoholic part formed is an enol, it would exist enolate under strongly alkaline condition. Assuming the anion in alkaline solution to be expressed approximately as VII, disappearance of optical activity would naturally result. Examination of the properties of Lactol secures structure VIII for this product and excludes another possibility such as tautomeric α,β-γ,δ-unsaturated carboxylic acid IX and α,β-unsaturated ketocarboxylic acid X. Thinking of the result on the titration of Lactol, carboxylic acid structures are clearly ruled out and the structure VIII is assigned to Lactol in solution. The use of ultraviolet spectrum is not reliable for the discrimination between lactol structure (VIII) and carboxylic acid structure X[18]; an absorption maximum in question is located near 245 mμ (Fig. 1), which would be reasonable both for Lactol (VIII) and α,β-unsaturated carboxylic acid (IX), considering the spectra of models G[19] and H[20].



λ_{max} 234 mμ λ_{max} 234 mμ

(G) (H)

From the position and the shape of hydroxyl stretching vibration which appeared in 3000 cm^{-1} region of the infrared spectrum, it becomes highly probable that there is no hydroxyl group of carboxylic nature. There was observed a band at 1708 cm^{-1}, the overlapping of which with a strong amide I band (1675 cm^{-1}) rendered the former to be a shoulder band of the latter. Utilizing this shoulder band, in conjunction with bands which appeared in sodium salt, it is possible to discriminate a structure VIII from IX and X. First, form IX is taken into account. Considering the infrared spectrum of model J, >C=O vibration of IX is expected to occur near 1660 cm^{-1}, which is not an actual case (band observed is at 1708 cm^{-1}).



$\nu_{C=O}$: 1665 cm^{-1}

(J)

Secondly, if Lactol took form X, it should, in addition to a band due to carboxylic acid, exhibit a band of isolated ketone at 1717 cm^{-1} (observed in sodium salt), which is absent in Lactol itself. Carboxylic acid structures are, in any event, inconsistent with the observation about ν_{O-H} region as described before.

Structure VIII can, in every respect, be re-

17) R= *p*-acetylamino-α-methyl cinnamylidene.

18) The possibility of IX is clearly excluded, since this should show an absorption at much longer wavelengths.

19) K. Bowden, E. A. Braude and E. R. H. Jones, *J. Chem. Soc.*, **1946**, 948.

20) E. R. H. Jones and M. C. Whiting, ibid., **1949**, 1419.

conciled with the infrared spectra of Lactol and of its sodium salt. Looking upon α,β-unsaturated lactone ($\nu_{C=O}$: 1717~1730 cm^{-1}) as the fundamental system, the structure VIII is constructed by attaching hydroxyl and methoxyl group to the fundamental system. These two substituents, in this case, cause opposite effects on the position of carbonyl absorption in the α,β-unsaturated ester system; i.e., carbonyl frequency tends to be lowered by the mesomeric effect due to the methoxyl group and to be raised by the inductive effect of the hydroxyl group. A slight lowering of carbonyl frequency would be expected as a net result, according to the estimation of two such effects based on the data[21] so far available. Thus, it is reasonable to ascribe a band at 1708 cm^{-1} to the lactonic carbonyl group in structure VIII. Furthermore, the appearance of a band at 1717 cm^{-1} in sodium salt is quite consonant with structure VIII.

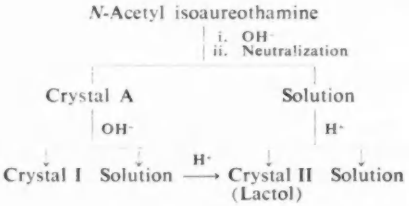
Chemical and spectral evidence now substantiates lactol structure VIII and rules out other possibilities, both in solution and solid state.

Experimental

Ozonization of Isoaureothin Derivatives.—In order to estimate the amount of methyl pyruvate produced by ozonization, the following procedure was applied in each run. Through a solution of ca. 500 mg. of sample in 10~15 ml. of chloroform, ozone was passed at −5~−10°C, until the uptake of ozone by sample ceased. On removal of the solvent at reduced pressure, a viscous syrup was obtained which was then treated at 60°C with 10~15 ml. of water in a flask equipped with a condenser. The aqueous fraction was separated, and a solution of 2,4-dinitrophenylhydrazine in 2 N hydrochloric acid was added to it. The crude hydrazone was collected, dried, dissolved in benzene and chromatographed on alumina. 2,4-Dinitrophenylhydrazone of methyl pyruvate, eluted with benzene-ether was collected and weighed, m. p. above 140°C, m. p. 145~147°C after recrystallization from benzene. Formaldehyde was also detected in every sample, except in the cases of *N*-acetyl aminol H (IIIb) and *N*-acetyl *O*-methyl aminol H (IIIe).

Lactol (VIII).—To a mixture of 110 ml. of ethanol and 55 ml. of 3 N aqueous sodium hydroxide, 2 g. of *N*-acetyl isoaureothamine was added. The flask was equipped with a reflux condenser and the suspension was warmed on a water bath at about 80°C for ten min., when the mixture became clear and of an orange color. After rapid cooling, the solution was neutralized with dilute hydrochloric acid; the solution turned yellow immediately at the neutralization point (pH 7). The solution was concentrated to 50 ml. under reduced pressure to yield fine crystals ("crystal A"). The yellow filtrate from "crystal A" was acidified gradually with dilute hydrochloric acid and the resultant precipitates (in part oily material) were collected and recrystallized from ethanol to yield almost colorless needles ("crystal II"), named Lactol, m. p. 170~171°C, 500 mg. When "crystal A" was dissolved in aqueous sodium hydroxide solution, a small part was insoluble in alkali. The alkali-insoluble crystals ("crystal I") were collected and recrystallized from ethanol repeatedly, m. p. 223~225°C, 40 mg.

The filtrate from "crystal I" was acidified and the resulting crystals were identified as "crystal II"; yield 250 mg., combined amount of "crystal II", 750 mg. The ultraviolet spectrum of "crystal II" showed λ_{max} 299 mμ (log ε 4.52) (in ethanol). $[\alpha]_D^{27}$, 0±1.5° (methanol, c=0.75). Titration was carried out in 66% aqueous ethanol solution.



Analysis was carried out on "crystal II" (Lactol):
Found: C, 67.06; H, 7.19; N, 3.63. Calcd. for $C_{24}H_{29}O_6N$: C, 67.43; H, 6.84; N, 3.28%.

Action of Diazomethane on Lactol.—Crystalline Lactol was added to ethereal diazomethane, but no reaction took place under the heterogeneous condition and the starting material was recovered. To 20 ml. of ethereal diazomethane was added 100 mg. of Lactol, dissolved in 20 ml. of methanol and the mixture was allowed to stand overnight. After removal of about two-thirds of the volume of solvent at reduced pressure, the precipitated crystals were filtered and recrystallized from ethanol to give colorless needles, m. p. 191~193°C, 25 mg., $[\alpha]_D^{28}$, 0±2° (c=0.8 in a mixture of chloroform and methanol (1:2)). The infrared spectrum taken in solid state (KBr) was identical with that of *N*-acetyl isoaureothamine (m. p. 209~211°C; $[\alpha]_D^{28}$ +140 ±1.5° in a mixture of chloroform and methanol (1:2)), the starting material of Lactol. Mixing two species of *N*-acetyl isoaureothamine (optical active and inactive species), no depression of the melting point was observed; mixted melting point, 193~200°C. Further concentration of filtrate gave a small amount of oily material, the infrared spectrum of which included bands at 1701(sh), 1690(s), 1606(m), 1592(m) and 1512(s) cm^{-1} (in chloroform). The main portion of the product might be an ester.

21) Literatures for the estimation of mesomeric effect of methoxyl group: (a) R. N. Jones and F. Herling, *J. Org. Chem.*, **19**, 1252 (1954); (b) R. S. Rasmussen and R. R. Brattain, *J. Am. Chem. Soc.*, **71**, 1073 (1949); (c) H. Heyman, S. S. Bhatnagar and L. F. Fieser, ibid., **76**, 3689 (1954).
Literatures as to inductive effect of hydroxyl (and alkoxyl) group: (d) R. B. Woodward and E. G. Kovach, *J. Am. Chem. Soc.*, **72**, 1009 (1950); (e) J. F. Grove and G. A. Willis, *J. Chem. Soc.*, **1951**, 877; (f) J. F. Grove ibid., **1952**, 3345; (g) L. A. Duncanson, J. F. Grove and J. Zealley, ibid., **1953**, 1331, 3637.

Action of Acetic Anhydride on Lactol.—To a mixture of 1 ml. of pyridine and 5 ml. of acetic anhydride was added 50 mg. of Lactol. It was kept at 70°C for ten minutes and after cooling, water was added to the solution, the resultant crystals were collected, recrystallized from ethanol and identified as the starting material, Lactol. Again, the same procedure was applied with prolonged heating (one hour); the resultant crystals, differing from Lactol, were recrystallized from ethanol and were identified as *N*-acetyl isoaureothamine by the comparison of the infrared spectra, m.p. 193~194°C.

Found: C, 70.59; H, 6.50. Calcd. for $C_{24}H_{27}O_5N$: C, 70.40; H, 6.50%.

All melting points were uncorrected. The ultraviolet spectra were measured with a Hitachi spectrophotometer and the infrared spectra with a Hilger H 800 spectrometer, equipped with rock-salt prisms. The optical rotations were measured with a Rudolf polarimeter.

Summary

During the course of the investigation of isoaureothin (I), several derivatives were obtained, the structure of which still remains unknown. The structure of these derivatives is presented in this paper on a basis of spectral and chemical evidence. The reduction of isoaureothamine (IIa) with lithium borohydride yielded an unexpected acidic substance, aminol H (IIIa) the formation of which, taken in conjunction with that of *N*-acetyl aminol A (IVa), provides an unusual type of reduction by complex metal hydride. Treatment of *N*-acetyl isoaureothamine (IIb) with alkali yielded an optically inactive substance, Lactol, which was formulated as VIII. The action of acetic anhydride on Lactol resulted in the formation of racemate of *N*-acetyl isoaureothamine (IIb).

This stuty was supported by a grant from the Ministry of Education, to which the authors are indebted.

Chemical Institute
Faculty of Science
Nagoya University
Chikusa-ku, Nagoya

Banyu Pharmaceutical Co.
Okazaki, Aichi-ken (T. N.)

SHORT COMMUNICATIONS

*NMR Spectra of the Metabolites of Monascus. I. NMR Spectra of Several Authentic Pyrones**

By Akira TERAHARA, Mamoru OHASHI, Koji NAKANISHI, Ichiro YAMAGUCHI and Naohiro HAYAKAWA

(Received July 15, 1960)

As with the C–2 and C–5 protons of furans[1] it may be expected that the protons alpha to the pyronoid oxygen will absorb at characteristically low frequencies and thus assist in the identification of the substitution pattern, which is not so facile by other physico-chemical methods. A Varian model V-4300 spectrometer was employedat 56.4 Mc; the band separations were estimated from the toluene peaks and chloroform was used as internal reference. The measurements were carried out in 10~20 percent solutions of 1 : 1 mixtures of chloroform and carbon tetrachloride. The results are summarized in Table I[2]. It can be seen qualitatively that the α-protons of both α- and γ-pyrones absorb at lower fields than the chloroform proton, and that they can be clearly differentiated from the other unsaturated protons. The coupling constant of the α- and β-protons, ca. 6 cps, is larger than the 1.3 cps value reported for furans[3]. The C–2 and C–6 methyl groups absorb in the range of 280~300 cps (II, III and VII). One of the nuclear protons of the extended pyronoid compound VII had its peak at an exceptionally low field, and this interesting phenomenon is presumably caused by the strong long range shielding effect of the nearby nitrile group. It is to be noted that this shift is opposite to that encountered in the well-known acetylenic protons[4] and this can be attributed to the directional difference of the triple bond[5]. Citrinin has been chosen

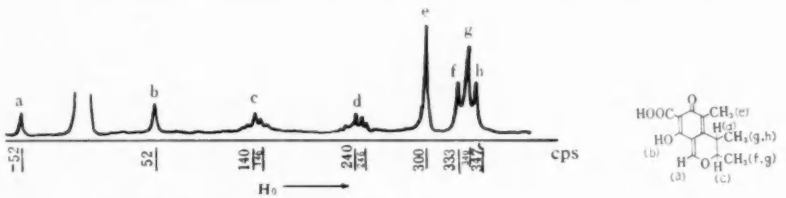
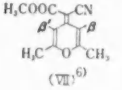


Fig. 1. NMR spectra of citrinin.

TABLE I. NMR SPECTRA OF PYRONES

Compound	α-H	β-H	Other H
γ-Pyrone(I)	−15, −9	71, 76	
2,6-Dimethyl-γ-pyrone(II)		73	289 (CH_3)
2-Methoxy-6-methyl-γ-pyrone(III)		87, 116	205 (OCH_3), 294 (CH_3)
2,6-Dimethoxycarbonyl-γ-pyrone(IV)		12	188 (OCH_3)
α-Pyrone(V)	−15	53, 60	
6-Methoxycarbonyl-α-pyrone(VI)		12	194 (OCH_3)
 (VII)[6]		−42, 52	200 (OCH_3), 284 (CH_3)

* Presented at the 13th Annual Meeting of the Chemical Society of Japan, Tokyo, April, 1960.

1) R. J. Abraham and H. J. Bernstein, *Can. J. Chem.*, 37, 1056 (1959).

2) No corrections for the bulk susceptibilities were carried out.

3) J. A. Pople, W. G. Schneider and H. J. Bernstein, "High-resolution Nuclear Magnetic Resonance", McGraw-Hill Book Company, New York (1959), p. 193.

4) We are grateful to Dr. N. Nakagawa, University of Electro-Communications, for this explanation.

5) L. M. Jackman, "Applications of NMR Spectroscopy in Organic Chemistry", Pergamon Press, London (1959), p. 114.

6) M. Ohta and H. Kato, This Bulletin, 32, 707 (1959).

as an appropriate model to clarify the spectra of the more complicated natural products, and its spectrum is analyzed in Fig. 1. The qualitative results obtained from the spectra of the pyrones are hoped to assist in handling related compounds.

We are grateful to Professor S. Fujiwara, the University of Tokyo for valuable suggestions.

Department of Chemistry
Faculty of Science
Tokyo University of Education
Bunkyo-ku, Tokyo
(A. T., M. O. & K. N.)

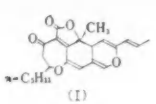
Division of Radiation Applications
Japan Atomic Energy Research Institute
Tokai, Ibaraki-ken
(I. Y. & N. H.)

*NMR Spectra of the Metabolites of Monascus. II. Monascorubrin**

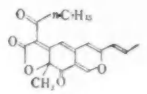
Mamoru OHASHI, Akira TERAHARA
and Koji NAKANISHI

(Received July 15, 1960)

Structures I[1] and II[2] have been tentatively proposed for monascorubrin[3], $C_{23}H_{26}O_5$, the orange pigment produced by *Monascus purpureus Wentii*. The NMR spectra of monascorubrin and dihydromonascorubrin have been measured in order to check the gross character of this complicated pigment, and if possible to differentiate between the two structures.



(I)



(II)

The measurements were carried out with a Varian V-4300 High Resolution Spectrometer operating at 40 Mc; chloroform was chosen as the inner reference and the band positions were read by the side-band method. It is seen in Fig. 1 that instead of the doublet "d" and "e" (propenyl methyl) and multiplet around "g", the spectrum of the dihydro-compound (Fig. 2) shows a triplet "j" which corresponds to the α-methylene of the propyl chain. Signals "a" and "b" are due to the *n*-alkyl chain and "c" arises from an uncoupled methyl group. Signal "i" corresponds in position to a proton adjacent to the pyronoid oxygen[4] and this is supported by chemical evidence[1]. The other two low-field signals, "g" and "h", undoubtedly arise from the remaining two nuclear protons. The shape and intensity of signal "f" is similar to "j" which has been assigned to a methylene group, and therefore in structure II it could be correlated with the $-CH_2CO-$ protons of the $n\text{-}C_7H_{15}CO-$ side-chain. On the other hand, in structure I the $-CH_2CO-$protons belong to the ABX type and thus a complicated pattern would be expected. The signal for the bridgehead proton of the ethereal ring in I can not be detected in the expected 100~200 cps region, but this could be due to extensive coupling with adjacent protons. The coupling of signals "h" and "i" (ca. 1.2 cps) presumably correspond to that observed between meta-oriented protons on benzene nuclei[5]. Thus, although no clear-cut differentiation between structures I and II is possible, the present measurements appear to be slightly in favor for the latter; however, chemical and ultraviolet spectroscopic evidence can be accounted for more smoothly

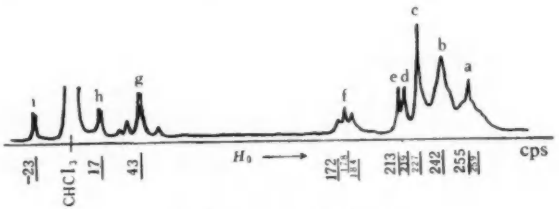


Fig. 1. Monascorubrin.

* Presented at the 13th Annual Meeting of the Chemical Society of Japan, Tokyo, April, 1960.

1) K. Nakanishi, M. Ohashi, S. Kumasaki and S. Yamamura, *J. Am. Chem. Soc.*, **81**, 6339 (1959).
2) B. C. Fielding, F. J. Haws, J. S. E. Holker, A. D. G. Powell, A. Robertson, D. N. Stanway and W. B. Whalley, *Tetrahedron Letters*, No. 5, 24 (1960).
3) A. Nishikawa, *J. Agr. Chem. Soc., Japan*, **8**, 1007 (1932).
4) Preceding paper.
5) J. A. Pople, W. G. Schneider and H. J. Bernstein, "High-resolution Nuclear Magnetic Resonance", McGraw-Hill Book Company, New York (1959), p. 193.

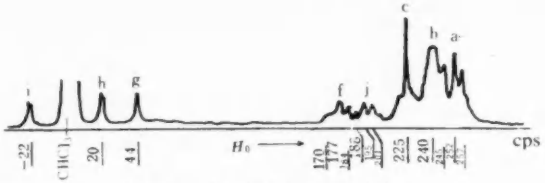


Fig. 2. Dihydromonascorubrin.

by structure I and this point will be reported shortly.

The authors are grateful to Professor Y. Yukawa, Osaka University, for measurement of the spectra and suggestions.

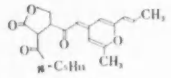
Department of Chemistry
Faculty of Science
Tokyo University of Education
Bunkyo-ku, Tokyo

*NMR Spectra of the Metabolites of Monascus. III. Monascoflavin and Reduction Products**

Mamoru OHASHI, Akira TERAHARA, Koji NAKANISHI, Ichiro YAMAGUCHI and Naohiro HAYAKAWA

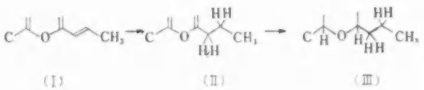
(Received July 15, 1960)

The structure of monascoflavin[1], $C_{21}H_{26}O_5$, m. p. 143°C, the yellow metabolite of *Monascus purpureus Wentii* has still not been elucidated. Its chemical properties showed that it was related to γ-pyrones[2], and the close relationship with monascorubrin has also been suggested[3]; structure I was provisionally proposed in 1959 as



(I)

a working hypothesis[4]. The NMR spectra of I, dihydro-I, hexahydro-I, and octahydro-I[5] have been measured with a Varian V-4300 model Spectrometer at 56.4 Mc, in order to correlate the chemical transformations with appropriate changes in the spectra (Figs. 1—4) and thus gain an insight into its structure. Chloroform was employed as an inner reference, and toluene was used for calculating line separation. Unlike monascorubrin[6], the present monascoflavin showed no peak in the −30 cps region corresponding to the α-hydrogen of γ-pyrones[7]. Instead of the "d" doublet (J: 6 cps, area for three protons) and the weak "k" multiplet in Fig. 1, the spectrum of II revealed a triplet "l" (J: 8 cps); this can be ascribed to the hydrogenation of the propenyl side-chain. The "a" and "b" peaks arise from the methyl and methylene groups, respectively, of the *n*-C_5H_{11} side-chain. The peaks at "c" is due to an uncoupled methyl group, and peaks "i" and "j" (Figs. 1 and 2) suggest that three protons attached to an unsaturated system are present besides the propenyl group. In Fig. 3, only a diffuse absorption "i" is left in this region; instead there appears two ill-defined peaks "m" and "n" which could be assigned to protons attached to α-carbons of an ethereal linkage. The hydrogenation stages can be represented by the following scheme:



(I) (II) (III)

Peaks "g" and "h" seem to originate from the γ-methylene of γ-lactone ring[8]; however, the coupling is not clear. The quadraplet "i" (Figs. 1 and 2) seems to arise from an AB type of coupling (J: 10 cps) and is absent in Fig. 3. Although it is apparent that the quadraplet is

* Presented at the 13th Annual Meeting of the Chemical Society of Japan, Tokyo, April, 1960.

1) A. Nishikawa, *J. Agr. Chem. Soc., Japan*, **8**, 1007 (1932); H. Salomon and P. Karrer, *Helv. Chim. Acta*, **15**, 18 (1931).

2) P. Karrer and A. Geiger, *Helv. Chim. Acta*, **25**, 289 (1941).

3) A. D. G. Powell, A. Robertson and W. B. Whalley, *Chem. Soc. Special Publ.*, No. 5, 27 (1956).

4) K. Nakanishi, M. Ohashi, A. Terahara, S. Kumasaki, S. Yamamura, Y. Hirata and H. Koike, the 3rd Symposium on the Organic Chemistry of Natural Products, Tokyo, 1959, Abstract, p. 55.

5) Chemical and spectroscopic results indicate that this is not a straight-forward reduction product. Accordingly, the spectrum will not be discussed at present.

6) Part II of this series.

7) Part I of this series.

8) L. M. Jackman, "Applications of NMR Spectroscopy in Organic Chemistry", Pergamon Press, London (1959), p. 55.

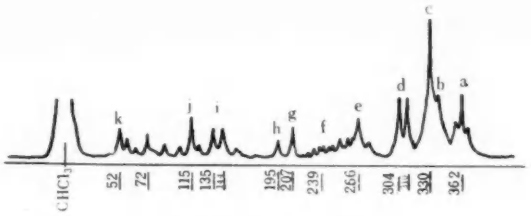


Fig. 1. Monascoflavin.

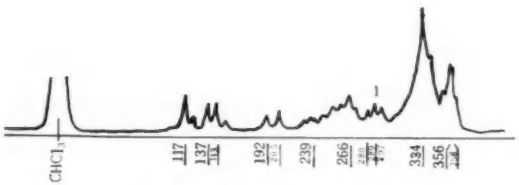


Fig. 2. Dihydromonascoflavin.

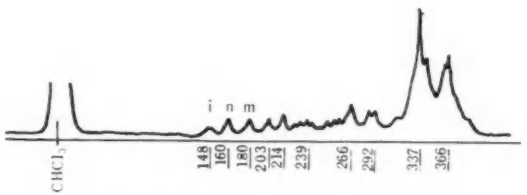


Fig. 3. Hexahydromonascoflavin.

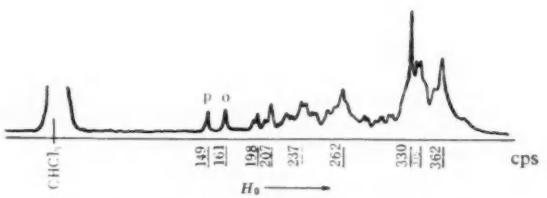
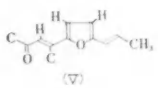


Fig. 4. "Octahydromonascoflavin".

due to protons on an unsaturated system, it is questionable that any two protons shown in the partial structure IV should couple so extensively.

(IV)



(V)

Structure V appears to be more appropriate to account for the coupling but this, in turn, can not be reconciled with ultraviolet spectral data[4]. Model compounds are being synthesized in order to check this point further. Peak "e" is presumably due to the $-COCH_2-$ protons but the reason that it is not split into a triplet is not clear. The diffuse "f" signal could arise from the bridgehead proton(s) of the β-keto-γ-lactone moiety.

Department of Chemistry
Faculty of Science
Tokyo University of Education
Bunkyo-ku, Tokyo
(M. O., A. T. & K. N.)

Division of Radiation Applications
Japan Atomic Energy Research Institute
Tokai, Ibaraki-ken
(I. Y. & N. H.)

On the Unusual Premelting Phenomenon of Potassium Thiocyanate Crystal

By Takako SHINODA, Hiroshi SUGA and Syûzô SEKI

(Received July 30, 1960)

Recently, Ubbelohde et al.[1] have published a series of papers on the unusual premelting effect of potassium thiocyanate crystal and concluded the rapid homophase increase of lattice defects as the melting point is approached. Quite independently,we have studied the transition phenomenon of this crystal in relation to that found in ammonium thiocyanate[2]. During the course of the differential thermal analysis of potassium thiocyanate we can not detect any such a tendency of premelting found by them, so we have reexamined the existence of the effect by use of the methods of dilatometric and d. c. conductivity measurements.

The sample (special grade, Wakō Pharmaceutical Co.) used were purified three times by fractional recrystallization from aqueous solution and then dried at 150°C under high vacuum (10^{-5} mmHg) for several hours after packing the sample in the glass-made dilatometer. Mercury was employed as a confining liquid, since it was confirmed that the mercury did not react with the sample even above the melting point. The results obtained are shown in Fig. 1 and in Table I in which the data by Ubbelohde and others are also given for comparison. For the d. c. conductivity measurement the specimen purified in the same way

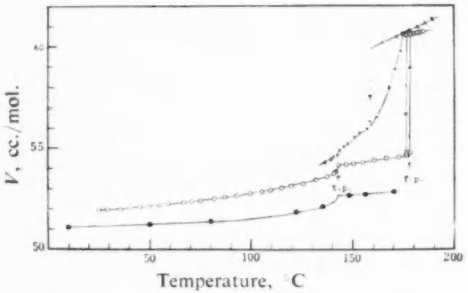


Fig. 1. The molar volume versus temperature curves of KSCN.
× Ubbelohde et al.
○ Present authors
● Yamada et al. (by X-ray method)

1) D. W. Plester, S. E. Rogers and A. R. Ubbelohde, *Proc. Roy. Soc.*, **A235**, 469 (1956); E. Rhodes and A. R. Ubbelohde, ibid., **A251**, 156 (1959); Idem, *Trans. Faraday Soc.*, **55**, 1705 (1959).

2) M. Sakiyama, H. Suga and S. Seki, presented at the 13th Annual Meeting of the Chemical Society of Japan, Tokyo, April, 1960.

TABLE I. VOLUME PARAMETERS OF KSCN

	Present authors	Ubbelohde and others
V_L	60.6	60.6±0.1
V_S	54.7	57.5
ΔV_f	5.9	3.1
$\Delta V_f/V_S$ (%)	10.8	5.4
V_t	54.13	54.7
$\Delta V_t/V_t$ (%)	1.53	0.36
ΔV_t	0.83	—
$V_{25°}$ (by flotation)	51.92*(51.3**)	51.92
t. p.	142.7°C	141.4°C
f. p.	178.45°C	175.1°C
α_S (high temp. phase)	2.6×10^{-4}	1.27×10^{-3}
α_L (low temp. phase)	3.91×10^{-4}	4.86×10^{-4}

V_L = molar volume of liquid at freezing point.
V_S = molar volume of solid at freezing point.
V_t = molar volume of solid at transition point.
ΔV_f = change in volume on melting.
ΔV_t = change in volume at transition point.
$V_{25°}$* = molar volume measured at 25°C by Ubbelohde et al.
** molar volume at ca. 10°C calculated from the dimensions of the unit cell[3].
(molar volumes are in cc./g. formula weight)

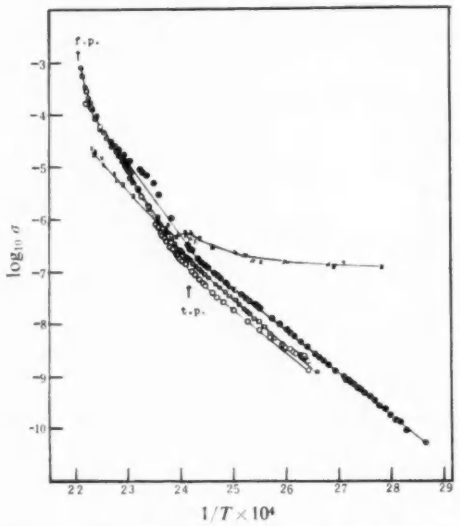


Fig. 2. The d. c. conductivity versus $1/T$ curves and their empirical equations of KSCN.
× Ubbelohde et al.
$\sigma_L = 1\times10^{-3}\times\exp(-6.6/\boldsymbol{RT})$
$\sigma_H = 3\times10^{14}\times\exp(-47.2/\boldsymbol{RT})$
● Our results
$\sigma_L = 2\times10^{12}\times\exp(-36.1/\boldsymbol{RT})$
$\sigma_H = 1\times10^{25}\times\exp(-59.5/\boldsymbol{RT})$
○ $\sigma_L = 3\times10^{10}\times\exp(-33.3/\boldsymbol{RT})$
$\sigma_H = 6\times10^{16}\times\exp(-44.0/\boldsymbol{RT})$
□ $\sigma_L = 2\times10^{11}\times\exp(-34.3/\boldsymbol{RT})$
$\sigma_H = 3\times10^{19}\times\exp(-49.2/\boldsymbol{RT})$
σ_L: d. c. conductivity for low temperature phase
σ_H: d. c. conductivity for high temperature phase

as mentioned above was used in the shape of disk tablet (weight : 100 mg. ; diameter : 13 mm. ; thickness : 1 mm.). The conductivity versus $1/T$ relationships and empirical equations derived are shown in Fig. 2 together with the results by Ubbelohde et al.

From the results of our dilatometric study it is found that the volume changes abruptly at the melting point with almost no hysteresis effect in the cooling direction, standing in remarkable contrast to the case by Ubbelohde et al., and that our data are in good harmony with the thermal expansion determined with the X-ray method by Yamada et al[3]. who kindly measured the temperature dependence of the unit cell dimensions in accordance with our request. The fact that we obtained much larger activation energy (~35 kcal./mol.) for the d. c. conductivity at lower temperature region seems to indicate that our sample involves less impurities and is also in parallel with the dilatometric data. Detailed descriptions of the experimental procedures and discussion will be published in due course.

Faculty of Science
Osaka University
Nakanoshima, Osaka

3) Y. Yamada, H. Suga and T. Watanabe, presented at the 7th Annual Meeting of the Society of Applied Physics, Japan, 1960.

Isotopic Exchange of Gallium between the Gallium(III) Ion and its Ethylenediamine-N, N, N', N'-tetraacetate and N-2-Hydroxyethylethylenediamine-N, N', N'-triacetate in an Aqueous Solution

By Kazuo Saito and Michiko Tsuchimoto (née Tamura)

(Received July 11, 1960)

As an extention of the present authors' study of the isotopic exchange of metal chelate compounds having d^{10} structure[1,2], they have examined the kinetics of exchange of gallium between the gallium(III) ion and its ethylenediamine-*N, N, N', N'*-tetraacetate (EDTA) and *N*-2-hydroxyethylethylenediamine-*N, N', N'*-triacetate (EDTA–OH) in water by the use of gallium-67. The radioisotope was produced by the irradiation of zinc with α-particles in a cyclotron and the γ-activity was measured with a scintillation counter having a sodium iodide crystal as scintillator. A gallium nitrate solution of a given pH was mixed with radioactive gallium EDTA or EDTA–OH complex solution of the same pH value and a part of the gallium ion was precipitated with 8-quinolinol at a certain lapse of time. The precipitate was filtered off, with a sintered glass filter plate, washed, dried, weighed and submitted to γ-ray counting.

The result is quite different from that for the corresponding exchange of indium[2]. The rate is very small in a neutral or an acid solution of pH more than 2, and is very great in a basic medium. In the pH range from 7 to 9, the exchange proceeds at a measurable rate. Since the exchange constant R expressed according to McKay's formula[3] is proportional to the concentration of the chelate b and is independent of that of the gallium ion a, the rate constant k is shown as R/b. Dependence of R upon various factors is shown in the Table together with other constants of the reaction. It is rather marked that the energy of activation E_A is very great as compared with that for the exchange of indium in an acid solution (10 kcal. per mol.). The present authors tend to suggest that the attack of the hydroxyl ion upon the central gallium in the co-ordination compound might play an important role in determining the exchange rate. Detail of the result and discussion on the reaction mechanism will be published later.

Table. Dependence of R upon various factors and some constants for the exchange in an aqueous solution of pH 7 to 9

	Ga-EDTA-OH	Ga-EDTA
$[OH^-]$	Proportional	Proportional
$[a]$	Independent	Independent
$[b]$	Proportional	Proportional
Ionic strength	Dependent to a less extent	Dependent to a less extent
Concn. of buffer	Almost independent	Almost independent
$k/[OH^-]_{25°C}$ (sec^{-1})	32	12
E_A (kcal./mol.)	27	25
A (sec^{-1})	3×10^{15}	2×10^{13}
ΔS (cal./mol./deg.)	−11	+1

At an ionic strength $\mu=0.2$ (KNO_3). A, frequency factor for k; ΔS, entropy of activation for k on the assumption $\kappa=1$ (κ, transmission coeff.), at pH 8.0.

Institute for Nuclear Study
The University of Tokyo
Tanashi, Tokyo

1) K. Saito and M. Tamura, This Bulletin, **32**, 533 (1959).
2) K. Saito and M. Tamura, *J. Inorg. Nur. Chem.*, **13**, 334 (1960).
3) H. A. C. McKay, *Nature*, **142**, 997 (1938).

The Polymerization of α-Methylstyrene by Iodine Catalyst

By Kozo Hirota, Gisuke Meshitsuka*, Fukuo Takemura and Tatsuto Tanaka

(Received May 19, 1960)

Although it is well known that iodine is an inhibitor for radical polymerization, its power of polymerizing some vinyl monomers has been reported by several investigators[1-4]. We have also found that iodine plays a role as an initiator for the polymerization of α-methylstyrene (αMS) at room temperature. Since it seems interesting that the reaction proceeds under the presence of lower concentration of iodine (0.1~0.3%) as compared with the case of polymerization of styrene by iodine (10~20%) recently reported[4], preliminary results are reported as follows.

It was found that the reaction proceeded even in the atmosphere but reproducible results were not obtained. The sample, therefore, was dehydrated in vacuo with metallic sodium or calcium hydride and was distilled into a reaction vessel without exposing to the air. The extent of polymerization of the monomer for a week reached maximum when the concentration of iodine added was ca. 0.1~0.3%. When more iodine than 1% was added to the monomer, the temperature of the reactant rose markedly high, as soon as iodine was allowed to contact wiih the monomer. In such cases the reactant became very viscous but the high polymers could not be obtained.

In Fig. 1, the yield of polymer for the reaction at 25°C vs. time are plotted. All the samples were enclosed at the same time and polymerized for different intervals of time, and the polymers were precipitated with excess of methanol. It is shown that the initial rate of polymerization in higher concentration of iodine (0.2%) is larger than that in lower one (0.1%), but in the former case the conversion did not become more than about 25%, while it increased up to more than 50% for an interval of 160 hr. in the latter case.

The color of iodine disappeared in both cases after a long period of reaction. The

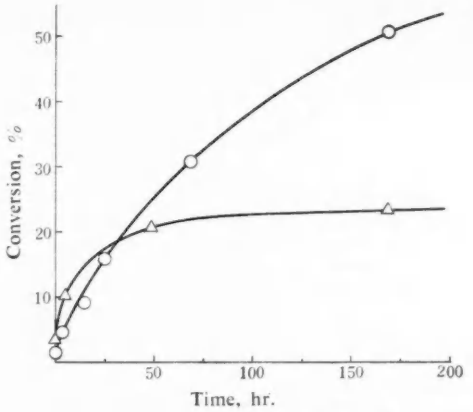


Fig. 1. Conversion of αMS vs. time.
Concn. of iodine: ○ 0.1%; △ 0.2%.

intrinsic viscosity of the polymers was independent of the time of polymerization, i.e., ca. 0.1** dl./g. and slightly larger in the case of lower concentration of iodine.

It is clear from the inhibitory power of iodine for polymerization by free radical mechanisms that this polymerization must be of non-radical and presumably of ionic character. It may be natural to assume that this reaction proceeds by a cationic mechanism as concluded in the polymerization of vinyloctylether[2] and of styrene[4] by iodine in view of the following experimental fact:—The polymer which is obtained from the mixture of αMS and methyl methacrylate (1:1 in volume ratio) in the presence of iodine showed essentially the same infrared absorption spectra as that of the polymer obtained from the similar mixture by stannic chloride.

Details of the experiment will be published in *J. Chem. Soc. Japan, Pure Chem. Sec. (Nippon Kagaku Zasshi)*.

Department of Chemistry
Faculty of Science
Osaka University
Nakanoshima, Osaka
(K. H., G. M. & T. T.)

Department of Chemistry
Faculty of Science
Nara Women's University
Nara (F. T.)

* Present address: Tokyo Metropolitan Isotope Center, Setagaya-ku, Tokyo.

1) W. Chalmer, *Can. J. Research*, **7**, 472 (1932).
2) D. D. Eley and A. W. Richard, *Trans. Faraday Soc.*, **45**, 425 (1949).
3) P. D. Bartlett and H. Kwart, *J. Am. Chem. Soc.*, **72**, 1056 (1950).
4) D. S. Trifan and P. D. Bartlett, ibid., **81**, 5573 (1959).

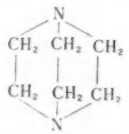
** This intrinsic viscosity corresponds to the molecular weight of ca. 3,000 (cf. C. P. Brown and A. R. Matheson, *J. Chem. Soc.*, **1958**, 3507).

Crystal Structure and Thermodynamical Investigations of Triethylenediamine, $C_6H_{12}N_2$

By Tadayoshi WADA*, Etsuko KISHIDA, Yujiro TOMIIE, Hiroshi SUGA, Syûzô SEKI and Isamu NITTA

(Received August 11, 1960)

Triethylenediamine, or 1, 4-diazabicyclo-(2, 2, 2)-octane, is considered to be a typical globular cage molecule with high symmetry. In order to amplify our knowledge on the thermal behavior of molecules in crystals, we have studied the crystal structure and some thermodynamical properties of this material.



The sample used for the present research was kindly offered us by Dr. Oae and by Houdry Process Corporation, Pa., U. S. A.

Small single crystals for the X-ray investigation, obtained from a saturated ethereal solution, were kept in glass capillaries to avoid humidity since the material has strong tendency of forming monohydrate. Useful X-ray data were obtained from Weissenberg photographs around a and c axes respectively taken with Ni-filtered Cu K_α radiation. The crystallographic data are: hexagonal system; $a=6.20\pm0.01$, $c=9.58\pm0.02$ Å; $Z=2$; $D_X=1.17$, $D_{obs}=1.14$ g. cm^{-3} [1]; space group $C_{6h}{}^2-P6_3/m$.

After seven and six successive calculations of the (F_o-F_c)-syntheses along [001] and [100] respectively, the final atomic parameters (Table I) and temperature factors were obtained. The interatomic distances thus calculated are given in Table II together with their standard deviations. Fig. 1 shows the molecular arrangement in a unit cell projected on (001).

TABLE I. ATOMIC PARAMETERS

	x	y	z
N	2/3	1/3	0.612
C	0.870	0.571	0.667
H_1	0.024	0.593	0.632
H_2	0.848	0.703	0.632

TABLE II. INTERATOMIC DISTANCES AND BOND ANGLES WITH STANDARD DEVIATIONS (Å UNIT)

C-C	1.590 (0.034)	C-C-N	110°53′
C-N	1.477 (0.017)	C-N-C	108°1′
C-H	0.96		
N···N	2.643 (0.020)	(in a molecule)	
	4.276, 6.932	(between molecules)	
C···N	3.748		
C···C	3.876, 3.977		

Molecules are approximately closest packed in the hexagonal lattice. The molecule in crystal has nearly $D_{3h}-\bar{6}m2$. The linkage of N-C-C-N shows a cis configuration. The crystal does not show any disordered structure in the room temperature phase.

In connection with the crystal analysis, differential thermal analysis was carried out up to 180°C. (Fig. 2). The sample, recrystallized from ethereal solution and sublimed repeatedly under high vacuum (10^{-5} mmHg), evacuated sufficiently at an elevated temperature prior to the measurement. The curve shows the presence of a phase transition at 79.8°C**, the entropy change of which is larger than that at the m. p. 161.1°C. These temperatures are more or less higher than those reported hitherto[1]. Measurements of vapor pressures have also been carried out below and above the transition point. The data are summarized in Table III.

The relatively small entropy of fusion shows that the high temperature phase of the crystal might belong to the category of the so-called

TABLE III. VAPOR PRESSURE DATA

Low temperature phase (51.3~78.6°C)

$\log p(\text{mmHg}) = -3233.6_3/T + 10.473_2$

$\Delta H = 14.8$ kcal./mol.

$\Delta S = 34.8$ e. u.

$\Delta G = 14.80_6 - 0.0347_6 T$ kcal./mol.

High temperature phase (80.0~96.3°C)

$\log p(\text{mmHg}) = -2722.1_3/T + 9.026_9$

$\Delta H = 12.5$ kcal./mol.

$\Delta S = 28.1$ e. u.

$\Delta G = 12.46_4 - 0.0281_4 T$ kcal./mol.

Transition (79.8°C)

$\Delta H = 2.3_4$ kcal./mol.

$\Delta S = 6.6_2$ e. u.

Fusion (161.1°C)

$\Delta H = 1.4_5$ kcal./mol.*

$\Delta S = 3.3$ e. u.

* This value was estimated from the differential thermal analysis curve.

* Present address: Yahata Iron and Steel Co., Ltd., Yahata.

1) A. Farkas, G. A. Mills, W. E. Erner and J. B. Maerker *I. Chem. Engineering Data*, **4**, 334 (1959).

** According to the private communication from Dr. E. F. Westrum the transition temperature is found to be about 353°K.

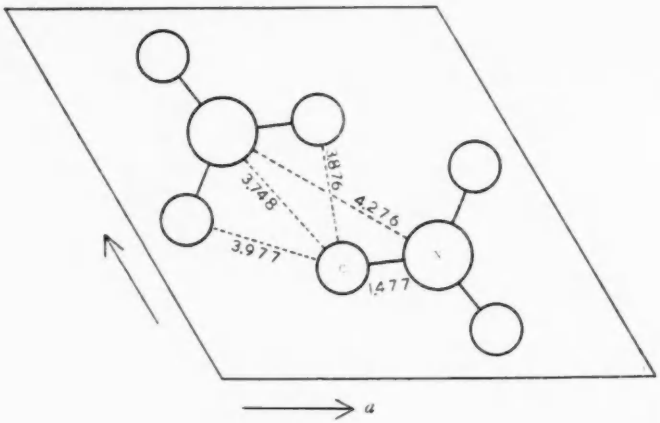


Fig. 1. The molecular arrangement in a unit cell projected on (001).

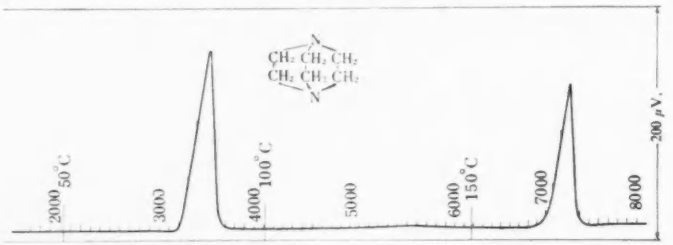


Fig. 2. Automatic recording curve of differential thermal analysis.

"plastic crystals" in the sense as suggested by Timmermans[2]. Accompanied with the large increase of entropy at the transition point, the molecule will gain some orientational, rotational or even translational freedom. Detailed analysis of the crystal structure in the high temperature phase is now going on. The heat capacities near the transition and melting points and the NMR measurement are currently under study in our laboratory. These studies may provide further informations on the mechanism of the phase transition and on the thermal motions in the high temperature phase.

Department of Chemistry
Faculty of Science
Osaka University
Nakanoshima, Osaka

2) J. Timmermans, *J. chim. phys.*, **35**, 331 (1938).